International Correspondence Schools
Scranton, Pa.

Inorganic Chemistry

Theory, Definitions, Laws, Nomenclature

PREPARED ESPECIALLY FOR HOME STUDY

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THEORY, DEFINITIONS, AND LAWS

GENERAL INTRODUCTION

1. Scope of Chemistry.—No science, perhaps, is of greater importance or broader in its scope than the science of chemistry. Together with physics it may be considered as the central science on which many other sciences, such as geology, biology, medicine, pharmacy, are based, or on which they depend for their fundamentals.

Chemistry touches all human interests. The multitude of familiar changes that take place about us daily, such as the rusting of iron, the decay of wood, the souring of milk, all come within its range of research and analysis. In fact, it is concerned with the identification, separation, and transformation of all matter—not only with matter in its natural states, but also with synthetic products, such as paints, explosives, plastics, and fertilizers.

Because of the diversified fund of information concerning the nature of matter and its changes which the science of chemistry has gathered and is still gathering, some knowledge of it is essential in all branches of engineering, in many of the professions, and in most manufacturing industries. The engineer who draws up specifications for a bridge or other structure, or the doctor who writes a prescription must first consider what properties are required by the aims in view, and then must ascertain what materials possess these properties. Much of his skill depends on his knowledge of chemistry.

2. Physics and Chemistry.—Like all other sciences, physics and chemistry treat of definite subjects. Physics is that branch of science which is concerned with changes in the form and state of matter, or in general, with changes that do not affect the composition of matter. Chemistry is that branch of science which deals with the composition of substances and with the changes which they may undergo. However, it should be borne in mind that chemical changes are always accompanied by physical ones, and that it is almost impossible to study chemistry without keeping in mind the physical principles involved.

3. Organic and Inorganic Chemistry.—Chemistry is a broad science, for it treats of every change in the universe in which the composition of matter is involved. In a general way it may be stated that the two classes of substances recognized are those that contain the element carbon and those that do not contain it. The part of chemistry that deals with substances which contain carbon is called organic chemistry, and the part that treats of substances which do not contain carbon is called inorganic chemistry. Some carbon compounds, such as carbon dioxide, carbon monoxide, carbonates, and carbon disulfide, are exceptions to these definitions and are described in both divisions of chemistry.

4. Development of Chemistry.—The development of chemistry from its obscure beginning to the present time, when it has become indispensable to mankind, is an interesting one. The desire to change common metals like lead into gold gave rise to a class of men known as alchemists. To them is due the foundation of what has since developed into a modern chemical science and industry, without which human progress would have been seriously retarded. It was not, however, until chemists began to study facts in relation to one another that certain fundamental laws were discovered and chemistry became a real science.

In early times it was supposed that earth, fire, water, and air were the only elements that constituted everything in the universe, and it was not until the latter part of the Eighteenth

Century that certain discoveries proved that water is composed of hydrogen and oxygen and that air consists chiefly of nitrogen and oxygen.

5. Priestley, an English clergyman, discovered oxygen. Scheele, a Swedish druggist, discovered chlorine. Cavendish, an Englishman, presented to chemistry a process that is today the basis of an important industry, the extraction of nitrogen from the air. He also showed that oxygen and nitrogen combine when an electric spark is passed through the air and that these elements, together with water, form nitric acid, a chemical which is indispensable in the manufacture of fertilizers and explosives. Lavoisier, a French chemist, was the first scientist to recognize clearly the importance of the discovery of oxygen; to him we owe our present knowledge of the relation of oxygen to combustion, or burning. In 1808, Dalton, an English school teacher, contributed to chemistry the atomic theory. Today this theory controls the activities of chemical enterprise and, though not yet entirely perfected, forms an indispensable part of chemical knowledge.

There were, of course, many other important discoveries that affected the progress of chemical knowledge greatly, but those mentioned are the fundamental ones upon which modern investigations and endeavors depend.

6. Recent Progress in Chemistry.—At present, chemistry has advanced to a position of prime importance and the theory of chemistry is so exact that one can now undertake various tasks with some measure of certainty and reasonable expectation of success. The opening of the European war in 1914 found America with certain great resources undeveloped and unused. Previous to that year it was cheaper to order from Europe whatever dyes, glassware, fertilizers, etc., were needed. The elimination of this source of supply meant that the natural resources of America had to be developed without delay.

The World War of 1941 found America in a similar position. Previous to this war, America could purchase raw silk and rubber from the Far East. The elimination of the supply of raw silk meant that the production of nylon had to be increased

in order to meet the demands for silk. Likewise, the elimination of the supply of raw rubber meant that the process of making synthetic rubber had to be developed without delay, for demands were urgent. Fortunately, chemistry had advanced sufficiently to overcome the many problems that arose during this period of development. Today, thanks to men who are chemically trained, American-made goods, such as potash products, glassware, porcelain ware, stoneware, coal-tar dyes, special metallurgical materials, drugs, photographic materials, plastics, and numerous synthetic products, can be obtained. Though some of the industries making these products may be only temporary in character, the wonderful development of them from mere ideas shows that chemical progress made since the days of the alchemist has placed chemistry in a position where it will always play an important role in supplying the needs and in conserving the resources of mankind.

COMPOSITION AND DIVISION OF MATTER

7. Matter.—Each of the terms used in chemistry has a distinct and definite meaning that must be understood by any one who wishes to study the laws and principles upon which the application of chemistry is based. Everything is subject to some kind of change that will alter its appearance, form, or chemical nature. Trees, grass, air, water—in fact, all substances—have two points of similarity. They all have weight and occupy space, and are called matter. Matter may be defined as anything that occupies space.

8. Forms of Matter.—A study of water will show that there are three physical states of matter: solids, liquids, and gases. When water is heated it changes to steam, and when it is cooled sufficiently it changes to ice. An analysis of steam, ice, and water will show that they all have the same definite chemical composition.

9. Properties of Matter.—A study of the different kinds of matter reveals the fact that each substance has characteristics or properties by which it can be recognized and identified. For example, one may recognize a substance as sugar if it is white

and sweet. Qualities by which matter can be identified are called properties of matter. Whiteness and sweetness are properties of sugar. In general, color, taste, and touch are some of the properties of matter that enable one to distinguish different substances.

The separate particles of crystallized sugar differ from those of powdered sugar in size and shape, but they resemble each other in color, taste, and touch. One can change crystallized sugar to powdered sugar by breaking up the former, thereby altering the size and shape of its particles. It is obvious that properties such as size and shape can be changed without destroying the nature of the substance, while other specific properties of a substance cannot be changed without altering its chemical composition. To change the color and sweetness of sugar, one must change the sugar to something that is not sugar. Properties that can be changed without altering the chemical nature of the substance are called conditional properties. Properties that cannot be changed without altering the chemical nature of the substance are called specific properties.

From the foregoing statements, it is apparent that matter has characteristics that affect the senses in different ways. The color, size, and shape of matter have definite effects on sight; its odor on smell; its sweetness, sourness, and bitterness on taste; and its stickiness and smoothness on touch. These characteristics are called properties of matter.

10. Divisibility of Matter.—The question naturally arises as to what constitutes a mass of matter—whether it is made up of particles, and, if so, how minute the divisions, or particles, of a mass can be made by dividing it by physical means, without destroying its specific properties. The view was once held that there was no limit, theoretically, beyond which matter could not be subdivided. This theory has been discarded for the theory which states that matter is composed of extremely small particles, and that if these small particles could be obtained by mechanical means, further subdivision by mechanical or physical means would be impossible. Furthermore, these minute particles, called molecules, are conceived as having the same

specific properties as the substances from which they are obtained. It must be remembered that the molecule is merely a conception, for it has never been isolated. A molecule may be defined as the smallest unit or particle into which a substance, an element, or a compound can be divided which retains all of the specific properties of that substance.

11. Hypothesis Defined.—The foregoing conception of molecules is known as the molecular hypothesis and is an excellent example of the line of reasoning often followed by scientists in their efforts to arrive at scientific truths. It is important, therefore, to understand at this point the meaning of the term *hypothesis*. Scientists note the various changes that take place, study them in detail, and seek to determine the causes of them. If no definite explanation results from their investigations, they proceed to imagine one, and then either prove or disprove it by tests. This imaginary explanation is called a hypothesis. The molecular hypothesis is accepted as true, for tests all seem to verify it and none seem to disprove it.

12. Molecules and Atoms.—Theoretically, but not actually, a mass of salt, for example, can be divided and subdivided repeatedly until an individual molecule of salt is obtained. A study of this molecule of salt shows that it has the same specific properties as the mass of salt; it has the same taste and no apparent odor. The molecule of salt differs from the mass only in conditional properties such as size and shape; chemically, both the mass and the molecule are the same, for the changes so far made are physical.

From what has already been said, it is obvious that the molecule is the smallest particle of a substance that can be obtained by physical means. There remains, however, chemical means by which division can be effected.

The atomic theory, first advanced by Dalton, involves the idea that molecules consist of atoms, particles that can be obtained by chemical means. A molecule of salt has been found by experiment to contain two simple substances, sodium and Na
Cl chlorine, in chemical combination. Both the mass and the mole-

NaCl

cule of salt have the chemical name sodium chloride. The elementary particles of sodium and chlorine, of which a molecule of salt is composed, are called atoms and can be obtained from a molecule of sodium chloride only by chemical means. An atom may be defined as the smallest particle of an element that is capable of entering into chemical combination.

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13. Mass and Weight.—The measure of quantity of matter that a body contains is the mass of that body and does not vary. Force The attraction that the earth has for a body is the weight of that body. The weight of a body is a variable factor that depends on the mass or quantity of matter that it contains, and its distance from the center of the earth. This textbook contains the same quantity of matter at any place on the earth or 1,000 miles above the earth. The mass is constant. However, as the distance of the book from the center of the earth varies, the attraction of the earth for it will also vary, thus causing the weight to change. (STANDARD WEIGHTS AT 760MM MERC. Hg

14. Substances.—It is necessary to apply the term *substance* properly, for it is used frequently in the study of chemistry. For example, sugar with its sweetness and whiteness is a substance, for it retains these specific properties no matter how much its conditional properties, size and shape, may vary.

Steel is a substance, for it retains its specific properties, hardness, strength, and color, no matter how much its conditional properties, shape and size, may be altered. It still remains the substance steel, even if made into a needle, a rail, or an axle.

Wood is a substance, since it retains its specific properties, its fibrous and fairly hard nature, no matter whether its form is that of a table, chair, or desk.

Matter, such as sugar, steel, or wood, retains its specific properties as long as its chemical composition remains the same. If one changes the chemical composition of sugar, the new product is no longer sugar, but a new substance possessing an entirely new set of specific properties. A substance may be defined as a homogeneous form of matter, that is, all parts are alike and show the same specific properties.

15. Elements and Compounds.—There are two kinds of molecules, those which contain similar atoms and those which contain dissimilar ones. Consider, for example, the substance mercuric oxide, each molecule of which contains one atom of mercury and one of oxygen. This substance, when heated, decomposes into the two substances, mercury and oxygen. There are no unlike atoms in either of these two substances. A molecule of oxygen contains nothing but oxygen atoms and a molecule of mercury contains nothing but a mercury atom. So far no one has been able to decompose either mercury or oxygen into simpler substances. An element may be defined as a substance that cannot be decomposed into simpler substances. Mercury and oxygen, for example, are elements. A compound is a substance that can be decomposed by chemical means into its constituent elements. Mercuric oxide is an example of a compound.

HgO

16. Physical Conditions of Elements.—Ordinarily, two of the elements, mercury and bromine, are liquids. Eleven of the elements, oxygen, hydrogen, chlorine, fluorine, nitrogen, argon, helium, krypton, neon, xenon, and radon, are gases; the rest of the elements are solids. Except carbon, the solid elements have been changed to liquids by means of heat. So far, carbon has been only slightly softened.

Hg - Br

O H Cl F N (argon) (helium) (krypton) (neon) (xenon) (radon)

The changing of a solid to a liquid is called fusing or melting; the changing of a liquid to a solid is called solidification; the changing of a gas to a liquid is called liquefaction; the changing of a solid or a liquid to a gas is called vaporization. As these terms are used constantly in chemistry, they should be thoroughly understood.

17. Chemical and Physical Changes.—There is a distinction between changes that are chemical and those which are physical. It is well to learn this distinction, for the whole of the science of chemistry is based on changes of some sort. If sugar is broken into smaller pieces, it still remains sugar, for only the conditional properties, shape and size, have been altered. If, on the other hand, intense heat is applied to the sugar, it soon darkens, loses its whiteness and sweetness, and

becomes something other than sugar, for the specific properties, taste and color, of the substance sugar have disappeared. In other words, the chemical composition of the substance sugar has been altered. The first change, in which sugar retains its chemical composition, is called a physical change, while the second change, in which the substance sugar is changed to an entirely different substance by altering its chemical composition, is called a chemical change. Hence, changes that do not affect the composition of substances are called physical changes, and changes in which substances disappear and something else is formed in their place are called chemical changes.

A lump of coal can be broken up and powdered finely without its color and hardness being destroyed; it still remains coal, even though changed in size and shape. This is a physical change. If, however, coal is burned, it loses its color and hardness, gives off a gas, and leaves a gray residue, called ash. The substance coal when burned loses its specific properties by which it is identified and forms new substances possessing other specific properties. This is a chemical change. It is apparent that the specific properties of a substance are destroyed in a chemical change and are not affected in a physical change. In other words, a substance is not changed chemically so long as it retains its specific properties, and it is changed chemically when its specific properties disappear.

18. The ultimate divisions of matter serve as the basis for simple definitions of chemical and physical changes. A physical change, as has been shown, can produce a molecule, but cannot affect its internal condition in any way, while a chemical change can act on a molecule to produce a change within it. Therefore, a physical change is one that does not affect a molecule and a chemical change is one that takes place within a molecule.

19. Mechanical Mixtures and Chemical Compounds.—It is necessary to know when the addition of one substance to another results in a chemical or in a physical change. This distinction can be made clear by considering the mixing of

sulfur, *S*, and iron, *Fe*. Sulfur* is a pale-yellow substance that melts easily when heated, and dissolves in carbon disulfide. Iron is a dark-gray metallic substance that is readily attracted by a magnet, and is not soluble in carbon disulfide.

When powdered sulfur and iron filings are rubbed together, the product, though differing in appearance from that of either of the constituents, is in reality the two original substances, sulfur and iron. The various particles of each are arranged side by side and can be separated from each other by mechanical means. The iron can be removed by passing a magnet over the mixture, or the sulfur can be freed from iron by treating the mixture with carbon disulfide, a process that dissolves the sulfur and leaves the iron unaffected.

If heat is applied to this mixture of sulfur and iron, a noticeable change takes place. The result is a black solid, which resembles neither iron nor sulfur and which is unaffected by a magnet or carbon disulfide. It is an entirely new substance known chemically as ferrous sulfide.

A mixture of sulfur and iron before the application of heat is called a mechanical mixture. A mechanical mixture may be defined as a collection of substances, either elements or compounds, that retain their specific properties and can be separated by physical means. A substance such as ferrous sulfide is called a chemical compound, which may be defined as a substance produced by the chemical reaction between two or more substances and which has specific properties different from the specific properties of any of the original substances.

20. A mechanical mixture of two substances is one in which the substances may be present in any proportions. It was thought at one time that, when two substances entered into chemical combination with each other, it was possible for them to do so in any proportion. Compounds would then be produced in which the composition would vary considerably. However, in 1806 this belief was found to be untrue. The theory was proved false when it was shown that, regardless of the source or

* *Sulfur* and *sulphur* are two ways of spelling the same word; *sulfur* is the preferable spelling.

the method of preparing a certain compound, the same constituents were always united in the same proportions by weight. Table salt, chemically known as sodium chloride, may be prepared by bringing together the elements sodium and chlorine. The same compound may also be prepared by placing sodium in contact with hydrochloric acid or by permitting hydrochloric acid to react with sodium carbonate. Regardless of the method of preparation, an analysis will invariably show that the elements, chlorine and sodium, have combined in the ratio of 1 part, by weight, of chlorine to .64859 part, by weight, of sodium. On a percentage basis it would appear as follows:

Sodium	39.35
Chlorine	60.65
	100.00

HCl+ → NaCl+

An analysis of sodium chloride dug from the earth or evaporated from sea water will show that the pure salt is constant in composition.

21. Dalton's Atomic Theory.—Dalton, an English school teacher, contributed to chemistry (1808) an atomic theory much like that of the early Greek philosophers. Like the Greek philosophers, he assumed that matter was composed of very small particles, which he also called atoms. Aside from this similar supposition, his views were quite different from those of the early philosophers. His theory embraced the added assumptions that atoms of the same element have the same weight, but atoms of different elements have different weights; and that the atoms of matter have the power of attracting each other and thereby effecting chemical union.

When Dalton announced his atomic theory, he failed to explain why some elements react with each other, while others do not. His theory also neglected to explain the forces holding the atoms together in molecules.

22. Modern Atomic Theory.—A discovery in 1876 led Sir William Crookes, an English physicist and chemist, to state that gases under greatly reduced pressure would conduct an

electric current and that fluorescence was produced in the gas, the color of which depended on the pressure. In the Crookes tube, Fig. 1, the positive terminal is called the anode, and the negative terminal is called the cathode. Both terminals are sealed in the tube. Crookes found that, when a charge of electricity was conveyed through the tube, the glow assumed a greenish color which appeared to be caused by rays emitted

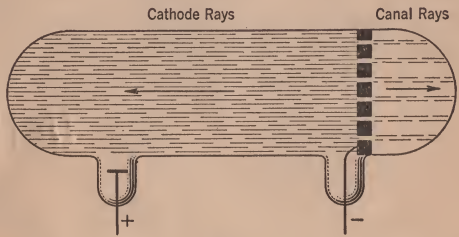


FIG. 1

from the cathode. A piece of metal placed in the path of these cathode rays became heated to incandescence and a magnet would deflect them. It appeared therefore, that they were composed of minute particles shot off from the cathode.

It can be shown that the cathode rays are capable of being bent under proper conditions. If a cathode tube is placed between two metallic plates, one carrying a high positive charge, and the other a high negative charge of electricity, the cathode rays will be bent toward the positively charged plate. As a result, it can be concluded that the ray emanation is of a negative nature and has its origin at the cathode regardless of the composition of the material; and since the same particles are thrown off from the cathode (or the gas), no matter what the cathode or gas, they must be constituent parts of all atoms.

In 1897, J. J. Thomson, a brilliant English physicist, discovered that all atoms are composed, in part, of these negatively charged particles. These particles are called electrons and have a mass $\frac{1}{1,850}$ as great as the mass of a hydrogen atom. Thomson's discovery of the electron was followed by Rutherford's discovery, in 1911, of the proton, a positively charged particle

that is a part of every atom. As a result of these discoveries, the modern atomic theory was propounded.

23. Scientists are now reasonably certain that the neutral atom is comparable to a miniature solar system. It is thought that the nucleus, or central portion, of the atom is composed of a number of electrons, or negative charges of electricity. In order to balance electrically the greater number of protons over the electrons, planetary electrons exist in the outer portion of the atom. It is thought that these outer electrons rotate rapidly about the nucleus. During the course of rotation, some of the outer electrons escape to other atoms. This is especially true with the more active metallic elements, such as potassium and sodium.

Contrasted with these are the non-metallic elements, such as chlorine and bromine, which exhibit a great tendency to accumulate electrons. It is now readily conceived why two metals such as potassium and sodium do not react with each other. The reason is also apparent why sodium and chlorine are ready to combine; the chlorine atom very readily accepts the electron which the sodium atom is so ready to give up. As a result of this exchange of electrons, a molecule of sodium chloride is formed and a chemical reaction has taken place. Picture in the mind's eye a sodium atom having a certain number of electrons and a like number of protons. If one of the planetary electrons is lost, there is an excess of protons over the electrons and the atom has a positive charge. A neutral chlorine atom has the same number of electrons as protons. However, if it takes on the electron that the sodium atom gives up, it immediately assumes a negative charge, since it has a greater number of electrons than protons. As a result, the two atoms, each having an opposite charge, are held together in the molecule by electrical attraction.

24. **Symbols.**—In indicating the composition of a substance, symbols are used instead of following the somewhat tedious practice of writing in full the names of the different elements involved. A symbol is usually the first letter of the name of the element capitalized. Examples are *O* for oxygen,

C for carbon, *N* for nitrogen, and *H* for hydrogen. When more than one element has the same initial letter, another letter that is in the name is added, but it is not capitalized. Examples are *Mn* for manganese, *Br* for bromine, and *Cl* for chlorine. In other cases, the symbol is taken from the Latin name. Thus, *Fe* is used for iron (*ferrum*), *Ag* for silver (*argentum*), and *Cu* for copper (*cuprum*). It is wise to learn the symbols of the different elements as we study them.

25. **Formulas.**—A combination of symbols is called a formula. The purpose of such a combination is to represent, in the briefest form possible, the elements of which a compound is composed. For example, a molecule of sodium chloride (common salt) consists of 1 atom of sodium and 1 atom of chlorine. The symbol of sodium (Latin, *natrum*) is *Na*; that of chlorine is *Cl*. Instead of stating that a molecule of sodium chloride consists of 1 atom of sodium and 1 atom of chlorine, the formula *NaCl* is used. From this formula a chemist can see at once the exact composition of this compound. Similarly, the formula *MgO* is a brief way of indicating magnesium oxide.

26. **Coefficient and Subscript.**—A number placed before a formula is called the coefficient. It shows the number of molecules represented. For example, *2KCl* represents 2 molecules of potassium chloride, and *3MgO* stands for 3 molecules of magnesium oxide. The numeral 1 is never used before a formula; if no number is given, 1 is understood.

The number placed to the right and a little below the symbol of an element is called the subscript. It indicates the number of atoms of that element contained in 1 molecule of the substance represented. For example, H_2 represents 1 molecule of the substance, or element, hydrogen, and the subscript 2 means that in this molecule there are 2 atoms of hydrogen.

In the formula of a compound, the number of atoms of each element present must be indicated. For example, the formula for water, H_2O, indicates that in each molecule of water there are 2 atoms of hydrogen and 1 atom of oxygen. In the formula $2K_2SO_4$, the coefficient shows that 2 molecules of the compound (potassium sulfate) are represented, while the subscripts show

that each molecule of the compound contains 2 atoms of potassium, 1 atom of sulfur, and 4 atoms of oxygen. It follows that the 2 molecules of potassium sulfate contain, in all, 4 atoms of potassium, 2 atoms of sulfur, and 8 atoms of oxygen. Thus, everything that comes after the coefficient is multiplied by that coefficient. K_2 S O_4

The formula $\boxed{Ca(OH)_2}$ represents 1 molecule of calcium chloride, which contains 1 atom of calcium, 2 atoms of oxygen and 2 atoms of hydrogen, the subscript 2 referring only to the contents of the parenthesis. The formula CaO_2H_2, though mathematically correct, is not used. $(Ca(OH)_2$

27. The foregoing principles are further illustrated by the following examples: NaI represents 1 molecuie of sodium iodide containing 1 atom of sodium and 1 atom of iodine. H_2S represents 1 molecule of hydrogen sulfide containing 2 atoms of hydrogen and 1 atom of sulfur. $\boxed{2KClO_3}$ represents 2 molecules of potassium chlorate, each molecule of which contains 1 atom of potassium, 1 atom of chlorine, and 3 atoms of oxygen; the 2 molecules contain, therefore, 2 atoms of potassium, 2 atoms of chlorine, and 6 atoms of oxygen. $2Ca(NO_3)_2$ represents 2 molecules of calcium nitrate, each molecule of which contains 1 atom of calcium, 2 atoms of nitrogen, and 6 atoms of oxygen; the 2 molecules contain 2 atoms of calcium, 4 atoms of nitrogen, and 12 atoms of oxygen.

After studying the different formulas presented, it may not be fully realized how it is known that, for example, sodium chloride contains 1 atom of sodium and 1 atom of chlorine. This perplexing problem may be answered by stating that the compound has been analyzed, and sodium and chlorine were found in the same proportion as indicated by their atomic weights.

28. **Order of Symbols.**—The order in which the symbols of the elements are placed in a formula for a compound depends upon the electrical properties of the element. These properties are designated as electropositive or electronegative, and for the present purpose it is necessary merely to say that the atom of the element is either positive or negative. However, it must

No

not be supposed that a sharp line can be drawn between the two classes of atoms.

In electricity it is a well-known fact that positive charges attract negative charges and that negative charges attract positive charges. It is true, also, that positive charges repel positive charges and negative charges repel negative charges. If these same principles are applied to positive and negative atoms, it will be seen that compounds are made up of positive and negative atoms. Some of these elements, according to their activity, are grouped in Table I.

TABLE I

ELECTROPOSITIVE AND ELECTRONEGATIVE SERIES

Positive Elements	Chemical Symbol	Negative Elements	Chemical Symbol
Lithium	*Li*	Fluorine	*F*
Potassium	*K*	Chlorine	*Cl*
Sodium	*Na*	Bromine	*Br*
Barium	*Ba*	Oxygen	*O*
Strontium	*Sr*	Iodine	*I*
Calcium	*Ca*	Sulfur	*S*
Magnesium	*Mg*	Phosphorus	*P*
Aluminum	*Al*	Nitrogen	*N*
Manganese	*Mn*	Carbon	*C*
Zinc	*Zn*	Silicon	*Si*
Chromium	*Cr*		
Iron	*Fe*		
Cadmium	*Cd*		
Cobalt	*Co*		
Nickel	*Ni*		
Tin	*Sn*		
Lead	*Pb*		
Hydrogen	*H*		
Antimony	*Sb*		
Bismuth	*Bi*		
Arsenic	*As*		
Copper	*Cu*		
Mercury	*Hg*		
Silver	*Ag*		
Platinum	*Pt*		
Gold	*Au*		

29. If now some of the more familiar formulas are examined it will be seen that most of them are made up of elements contained in Table I, and that, in constructing them, the positive element is written first. Consider the formulas for sodium chloride, $NaCl$; potassium nitrate, KNO_3; potassium chloride, KCl; sodium iodide, NaI; hydrogen sulfide, H_2S; and calcium nitrate, $Ca(NO_3)_2$. In $NaCl$ the positive Na is written first and the negative Cl last; in KNO_3, the positive K is written first, the negative N second, and the still more negative O last; in $Ca(NO_3)_2$, the positive Ca is written first, the negative N second, and the still more negative O, last. In the remaining formulas, KCl, H_2S, and NaI, it will be found upon inspection that the statements made concerning the order of symbols hold good. When two symbols are enclosed in parentheses, they represent two elements which act as one and the enclosed portion of the formula is called a radical. Thus, (OH) and (NO_3) are known as the hydroxyl and the nitrate radical, respectively. In sulfuric acid, H_2SO_4, the SO_4 group is known as the sulfate radical. In phosphoric acid, H_3PO_4, the PO_4 group is called the phosphate radical.

GASES, LIQUIDS, AND SOLIDS

METRIC WEIGHTS AND MEASURES

30. **Fundamental Units.**—The metric system is commonly used by chemists for expressing quantities of weight, volume, and distance.

The fundamental unit of length in this system is the meter, which is equivalent to 39.37 inches. From it are derived the units of capacity and weight, called the liter and gram, respectively. All other units of the metric system are decimal subdivisions and multiples of these fundamental units; that is, they are tenths, hundredths, etc., and ten times, one hundred times, etc., the meter, liter, and gram. The prefixes that are used in the metric system are shown in Table II.

The liter and gram bear a simple relation to each other; for instance, 1,000 milliliters are equal in capacity to 1 liter, and 1 liter of pure water at 4° C. weighs 1,000 grams.

TABLE II

PREFIXES OF THE METRIC-SYSTEM UNITS

Prefixes	Meaning of Prefixes	Names of Units
milli-	one-thousandth $=\frac{1}{1,000}=.001$	meter (for length)
centi-	one-hundredth $=\frac{1}{100}=.01$	gram (for weight)
deci-	one-tenth $=\frac{1}{10}=.1$	liter (for capacity)
unit-	one, or 1	
deka-	ten, or 10	
hecto-	one hundred, or 100	
kilo-	one thousand, or 1,000	

31. **Metric System.**—Basically, the metric system is a decimal system that makes calculations much easier than the method which uses the English equivalents. In Table III are shown metric units with the corresponding English system equivalents.

TABLE III

METRIC SYSTEM OF MEASUREMENT

Unit	Abbreviation	Metric Equivalent	English System Equivalent
Length			
Millimeter	mm.		.03937 inch
Centimeter	cm.	10 mm.	.3937 inch
Meter	m.	100 cm.	39.37 inches
Kilometer	km.	1,000 m.	.62 mile
Volume			
Cubic centimeter	c.c.		.06 cu. in.
Milliliter	ml.	1.000027 c.c.	
Liter	l.	1,000 ml.	61.03 cu. in.
Weight			
Milligram	mg.		.0154 grain
Gram	g.	1,000 mg.	.035 oz. avoir.
Kilogram	kg.	1,000 g.	2.205 lb. avoir.

TABLE IV

METRIC UNITS AND CONVERSIONS

To Convert From	To	Multiply by
Pounds, avoirdupois	kilograms	.45359
Ounces, avoirdupois	grams	28.3495
Kilograms	pounds, avoirdupois	2.2046
Grams	ounces, avoirdupois	.03527
Pounds, troy	kilograms	.37324
Kilograms	pounds, troy	2.6792
Ounces, troy	grams	31.1034
Grams	ounces, troy	.03215
Grams	grains	15.4324
Grains	grams	.06479
Pounds per cu. in.	grams per c.c.	27.6800
Pounds per cu. ft.	kilograms per cubic meter	16.018
Grams per c.c.	pounds per cu. in.	.03613
Milligrams per liter	parts per million	1.0000
Grains per gal. (U. S.)	parts per million	17.118
Parts per million	grains per gal. (U. S.)	.0584
Gallons	liters	3.7853
Liters	gallons	.26417
Liters	quarts (liquid)	1.0567
Quarts (liquid)	liters	.94633
Liters	quarts (dry)	.90809
Quarts (dry)	liters	1.1012
Ounces (fluid)	cubic centimeters	29.5737
Cubic centimeters	ounces (fluid)	.03381
Meters	yards	1.0936
Yards	meters	.91440
Centimeters	inches	.39370
Inches	centimeters	2.5400
Miles	kilometers	1.6093
Cubic feet	liters	28.316

32. Common Metric Units in Use.—Practically all of the weights and measures of the metric system that are commonly used by the chemist are shown in Table IV.

MEASUREMENT OF TEMPERATURE

33. Thermometers.—Chemists use instruments to measure the intensity of heat, that is, temperature, as heat has a strong influence on chemical action. Based on the fact that heat causes

substances to expand, an instrument called a thermometer is constructed to measure temperatures. The substance used in it, usually mercury, is sensitive to heat changes and thus affords a means of determining the temperature. The mercury in the tube of the thermometer, Fig. 2, expands and rises as the temperature increases, and contracts and falls as the temperature decreases. While thermometers differ in construction and materials, they are based on the same principle.

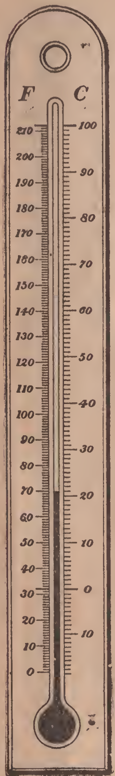


FIG. 2

34. **Temperature Scales.** — Thermometers are marked off into divisions, as shown in Fig. 2, each of which is called a degree and is expressed as 1°. The ordinary house thermometer, generally used in the United States, is based on the Fahrenheit scale. On this scale, the freezing point of pure water is called 32° F. and the boiling point 212° F. The space between 32° F. and 212° F. is divided into 180 equal parts, each one of which is called a degree Fahrenheit, expressed as 1° F. The scale is usually extended below and above 32° F. and 212° F., respectively, the same divisions being used.

On the centigrade scale, generally used by scientists, the freezing point of water is called 0° C. and the boiling point 100° C. The divisions representing degrees are secured by dividing the space between 0° C. and 100° C. into 100 equal parts. These divisions are also extended below 0° C. and above 100° C.

35. It is frequently necessary to change from one scale to the other. To change from Fahrenheit to centigrade, 32 is subtracted from the given temperature and the result is multiplied by five and divided by 9. This process may be expressed by the formula

$$C=\frac{5}{9}(F-32) \tag{1}$$

in which C is centigrade temperature and F is Fahrenheit temperature.

To change from centigrade to Fahrenheit, multiply the given temperature by 9, divide the result by 5, and add 32. This statement may be expressed by the formula

$$F=\frac{9}{5}C+32 \tag{2}$$

EXAMPLE 1.—Convert 41° F. to the corresponding centigrade temperature.

SOLUTION.—Apply formula 1. Substituting the known values gives

$$C=\frac{5}{9}(41-32)$$

Then, $$C=\frac{5}{9}(9)$$

or $C=5°$. Ans.

EXAMPLE 2.—Convert 10° C. to its corresponding Fahrenheit temperature.

SOLUTION.—Apply formula 2. Substituting the known values gives

$$F=\frac{9}{5}(10)+32$$

Then, $F=18+32$

or $F=50°$. Ans.

36. **Absolute Temperature and Charles' Law.**—During the latter part of the Eighteenth Century and the early part of the Nineteenth Century, considerable work on the influence of temperature changes on the behavior of gases was done by different investigators, prominent among whom were Gay-Lussac and Charles. It was Charles who, in 1787, made the discovery that when the pressure remains constant the volume of any gas is directly proportional to the absolute temperature. This is known as Charles' law which may be interpreted to mean that, given a known volume of a gas at a certain temperature and pressure, the volume of this gas will increase or decrease a like amount for every degree increase or decrease in its temperature, provided the pressure upon it is kept constant. The extent to which a gas expands for each degree rise in temperature is called the coefficient of expansion and is represented by the figure .003676, or approximately $\frac{1}{273}$. Thus, if a liter of oxygen

at 0° C. and 760 millimeters pressure is heated to 1° C. and the pressure is readjusted to 760 millimeters, the volume will have increased $\frac{1}{273}$ and will occupy a space of $1\frac{1}{273}$ liters.

On the other hand, a gas will also contract $\frac{1}{273}$ of its volume for every degree decrease in its temperature, so that when a point 273° below zero centigrade is reached, the gas, according to Charles' law, would occupy no volume. This point is known as absolute zero. However, the conception would hardly agree with the theory regarding the conservation of matter, so in this case the space between the molecules of the gas must be taken into account. Theoretically, when a gas is cooled, the intermolecular spaces are gradually decreased until at −273° C. they no longer exist. Actually, however, the gas would liquefy before a temperature of −273° C. was reached and Charles' law could not be applied.

Facts such as these are easily verified by experiment and may be expressed by the formula

$$\frac{V}{V'}=\frac{T}{T'}$$

in which V =original volume;
V'=new volume;
T =original absolute temperature;
T'=new absolute temperature.

The temperatures represented by T and T' are measured on the absolute scale and are therefore equivalent to 273 plus the temperature on the centigrade scale. For instance, 10° on the centigrade scale is equivalent to 273+10, or 283° absolute, and, for the sake of clearness, the formula may be written

$$\frac{V}{V'}=\frac{T+t}{T'+t'}$$

in which V =original volume;
V'=new volume;
T =273°;
T'=273°;
t =original degrees centigrade;
t'=new degrees centigrade.

Several examples showing how the different values are substituted for the letters of the formula will make this phase of the subject more clear.

EXAMPLE 1.—If a volume of oxygen of 100 liters at 12° C. is cooled down to a temperature of 5° C., what will be the new volume occupied by the gas?

SOLUTION.—By substituting the known values in the formula,

$$\frac{100}{V'}=\frac{273+12}{273+5} \text{ or, } \frac{100}{V'}=\frac{285}{278}$$

Then, $285V'=278\times 100$

Therefore, $V'=\frac{278\times 100}{285}=97.54$ liters, new volume.

EXAMPLE 2.—A gas occupies a volume of 400 milliliters at 35° C. What volume will it occupy when the temperature is lowered to −15° C.?

SOLUTION.—By substituting the known values in the formula,

$$\frac{400}{V'}=\frac{273+35}{273-15} \text{ or, } \frac{400}{V'}=\frac{308}{258}$$

Then, $308V'=258\times 400$

Therefore, $V'=\frac{258\times 400}{308}=335.06$ milliliters, new volume. Ans.

MEASUREMENT OF PRESSURE

37. Barometers.—When the study of the properties of gases is taken up, pressures are used in the different problems involved. Hence, for the sake of clearness, one type of instrument by which pressures are measured, called a mercurial barometer, is here described.

With the mercurial barometer, the pressure is indicated by the height of a column of mercury, and the pressure is usually expressed in either inches or millimeters. In Fig. 3 the space *a* is almost completely exhausted or freed from air and vapor, so that the column of mercury *b* will not be depressed. The space *c* has access to the atmosphere through a small opening *d*, and the pressure in this space is the same as that outside the tube. The pressure in space *c* forces column *e* down and column *b* up. Manufacturers of these instruments graduate them either in inches or in millimeters, so that they record pressures accu-

rately. Normal pressure of the atmosphere at sea level and 0° C. is equal to the pressure exerted by a column of mercury 760 millimeters, or 29.922 inches, high. In chemistry, pressure is generally expressed in millimeters.

38. Ways of Expressing Pressure.—Several ways of expressing pressure are in use. One is based on the number of pounds exerted on a square inch of surface; another is based on the number of kilograms exerted on a square centimeter of surface; another on the height of a column of mercury, either in inches or millimeters; and still another on the height of a column of water, either in feet or inches. Table V can be used

TABLE V

PRESSURE EQUIVALENTS

	Pounds per Sq. In.	Inches of Mercury	Mm. of Mercury	Feet of Water	Inches of Water	Kg. per Sq. Cm.
Pounds per sq. in.	1.000	2.0360	51.715	2.3066	27.679	.070307
Inches of mercury	.49116	1.000	25.4	1.13299	13.596	.03453
Mm. of mercury	.019337	.03937	1.000	.0445	.5340	.0013595
Inches of water	.036136	.07355	1.868	.0834	1.000	.0025399
Kgm. per sq. cm.	14.223	29.0267	735.56	32.845	394.140	1.000

to change the pressure expressed in terms of one system to those of another. For example, a pressure of 5 pounds per square inch is equivalent to 5 times 2.0360, or 10.18 inches of mercury; likewise 20 millimeters of mercury is equivalent to 10.68 inches of water.

The phrase, *under normal conditions of temperature and pressure,* which is frequently met in studying properties of gases, means 0° C. and 760 mm. As previously stated, a pressure of 760 mm. is equivalent to normal, or standard, air pressure at sea level, which in turn is equivalent to 14.7 pounds per square inch. This pressure is usually designated as 1 atmosphere.

39. **Boyle's Law.**—The effect of pressure on the volume of gases was investigated by Robert Boyle. In the year 1660 he discovered that, if the temperature remains constant, the volume of any gas varies inversely as the pressure. In other words, the volume of a gas decreases as the pressure upon it increases, and

FIG. 3

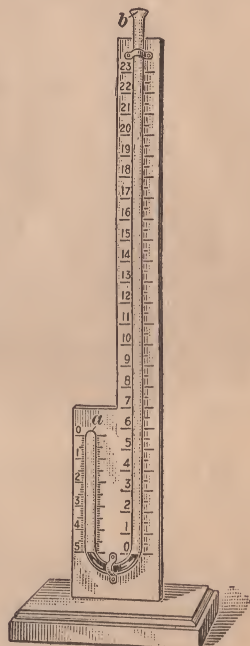


FIG. 4

increases as the pressure is decreased. This discovery is known as Boyle's law and may be expressed by the formula

$$\frac{V}{V'} = \frac{P'}{P}$$

in which V = original volume;
V' = new volume;
P' = new pressure;
P = original pressure.

That Boyle's law holds good for all practical purposes may easily be shown by a simple experiment. The apparatus shown

in Fig. 4 consists of a glass U-tube mounted on a wooden stand. The limb *a* of the tube is sealed and the upper end of the limb *b* is open to the air. If a quantity of mercury is poured into the open end so as to fill the bend of the tube completely as shown, and if then the tube is tilted, if necessary, to allow some of the air to escape from the short limb, the levels of the mercury in each limb are made equal. The pressure on the mercury in the limb *b* will be the same as that of the atmosphere, and the pressure on the confined air in the closed limb *a* will be 1 atmosphere. If sufficient mercury is added so that a column in the limb *b* of about 30 inches is obtained, a pressure of 2 atmospheres will be exerted on the air in the limb *a*, and the volume of this air will be found to have decreased one-half. It should be remembered that 29.922 inches, or 760 millimeters of mercury, is equivalent to a pressure of 1 atmosphere. With apparatus sufficiently large and strong, it can be shown that a given volume of gas can be made to decrease according to the increase in pressure. Thus, a volume of 1 liter is decreased to $\frac{1}{2}$ liter by a pressure of 2 atmospheres; to $\frac{1}{3}$ liter by a pressume of 3 atmospheres, and so on.

The use of the formula expressing Boyle's law in solving practical problems is shown by the following examples:

EXAMPLE 1.—A gas occupies a volume of 100 liters at a pressure of 725 millimeters. What volume will it occupy at 760 millimeters?

SOLUTION.—Substituting the known values in the formula gives

$$\frac{100}{V'}=\frac{760}{725}$$

Then, $760V'=725\times100$

Therefore, $V'=\frac{725\times100}{760}=95.39$ liters, new volume. Ans.

EXAMPLE 2.—A gas occupies a volume of 10 milliliters at a pressure of 780 millimeters. What will be the volume if the pressure is decreased to 750 millimeters?

SOLUTION.—Substituting the known values in the formula gives

$$\frac{10}{V'}=\frac{750}{780}$$

Then, $750V'=780\times10$

or $$V' = \frac{780 \times 10}{750}$$

Therefore, $V' = 10.4$ ml., the new volume. Ans.

40. Combined Effects of Temperature and Pressure.—In general practice, the chemist usually must make corrections for both pressure and temperature in reducing the volume of a gas to standard conditions. It is, therefore, helpful to combine the laws of Boyle and Charles into one equation, which may be expressed by the following formula:

$$V' = \frac{VT'P}{TP'}$$

This combined formula is derived as follows: According to Charles' Law,

$$\frac{V}{V'} = \frac{T}{T'}$$

or $$V' = \frac{T'V}{T}$$

Also, according to Boyle's law,

$$\frac{V}{V'} = \frac{P'}{P}$$

In this formula, V, the original volume, is the same as V', the new volume in the formula for Charles' law. If the value of V' is substituted for V in the formula for Boyle's law the following formula is obtained:

$$\frac{\frac{T'V}{T}}{V'} = \frac{P'}{P}$$

Simplifying, $$\frac{T'V}{T} = \frac{V'P'}{P}$$

or $$TV'P' = VT'P$$

Therefore, $$V' = \frac{VT'P}{TP'}$$

in which V' = new volume;
V = original volume;
P' = new pressure;
P = original pressure;
T' = new absolute temperature;
T = original absolute temperature.

It must be remembered that temperatures are measured on the absolute scale and that both T' and T represent 273+ degrees centigrade. The following example illustrates the application of the formula:

EXAMPLE.—A gas occupies a volume of 250 liters at a pressure of 745 millimeters and at a temperature of $-15°$ C. What volume will it occupy if the pressure is increased to 765 millimeters and the temperature is raised to 20° C.?

SOLUTION.—Substituting the known values in the formula gives

$$V' = \frac{250 \times (273+20) \times 745}{(273-15) \times 765}$$

Therefore, $V' = 276.4$ liters. Ans.

41. Dalton's Law.—Each gas in a mixture of gases exerts its own particular characteristics just as if it were the only gas present. For instance, oxygen in the atmosphere acts as if it were alone, exerting a pressure equal to only one-fifth of the atmospheric pressure, because it occupies only one-fifth of the volume of the atmosphere.

As a result of the investigation of many mixtures of gases, Dalton, in 1807, formulated a law which may be stated as follows: *The pressure of a mixture of gases is equal to the sum of the pressures of the components, and the pressure of each component of a mixture depends on the concentration of that component.* The law may be applied in cases where gases are confined over water. The water vapor diffuses into the gas according to the temperature and pressure surrounding it, so that when the volume of gas is measured it consists of both the gas and the water vapor. The water vapor, being a gas, exerts its own individual pressure and must be considered in any subsequent calculations. The pressure exerted by the vapor is known as vapor tension. In Table VI are shown different pressures exerted

by water vapor at different temperatures. If the pressure of a certain volume of gas in a confining space over water at 30° C. is found to be 750 mm., the actual pressure cannot be attributed to the gas alone, for the vapor tension is partly responsible.

TABLE VI

VAPOR TENSION OF WATER

Temperature Degrees C.	Pressure Mm. of Mercury	Temperature Degrees C.	Pressure Mm. of Mercury
0	4.6	24	22.4
5	6.5	25	23.8
8	8.0	26	25.2
9	8.6	27	26.7
10	9.2	28	28.3
11	9.8	29	30.0
12	10.5	30	31.8
13	11.2	35	42.1
14	12.0	40	55.3
15	12.8	50	92.5
16	13.6	60	149.4
17	14.5	70	233.7
18	15.5	80	355.1
19	16.5	90	525.8
20	17.5	100	760.0
21	18.7	150	3581.0
22	19.8	200	11688.0
23	21.1	230	20925.0

From Table VI it is found that at 30° C. the tension of water vapor is 31.8 mm. The true pressure due to the gas alone is $750-31.8=718.2$ mm.

In the calculation of gaseous volumes confined over water, the vapor tension must always be considered.

EXAMPLE.—A certain gas, confined over water, measures 300 ml. at 25° centigrade and 770 mm. of pressure. What will be the volume of the dry gas under normal, or standard, conditions?

SOLUTION.—From Table VI it is found that at 25° C. the vapor tension of water exerts a pressure of 23.8 mm. The true pressure due to the gas alone is then $(770-23.8)$. As previously stated, normal conditions

means 0° C. and 760 mm. Substituting the known values in the formula gives

$$V' = \frac{300 \times 273 \times (770 - 23.8)}{298 \times 760}$$

Therefore, $V' = 268.2$ ml. Ans.

DENSITY OF GASES

42. Diffusion.—As the molecules of all substances are in a constant state of motion, it is natural to suppose that, if two different non-reacting substances in a pure state are brought together, their molecules will intermingle. That this is exactly what occurs, may be easily proved, and although, in the case of solids, the process takes place at a much slower rate than with liquids or gases, this is what might be expected according to the kinetic theory. But the molecules of solids are still in motion to the extent that carbon is known to mix with pure hot iron, and gold with lead. The time occupied before an appreciable change takes place is in most cases a matter of years.

With liquids of mutual solubility, mixing, or diffusion, takes place much more rapidly; thus, if glycerin is placed in contact with water, a solution will soon result. It is in the case of gases that the phenomenon of diffusion is most striking. It may be shown that, regardless of their specific gravities or densities, gases will mix. If the mouth of a 1-liter bottle containing hydrogen is inverted over the mouth of a 2-liter bottle containing carbon dioxide, part of the hydrogen will pass into the bottle containing carbon dioxide, and part of the carbon dioxide will pass into the bottle containing hydrogen. Finally, each bottle will contain 1 part, by volume, of hydrogen to 2 parts, by volume, of carbon dioxide. The 1-liter bottle will contain $\frac{1}{3}$ liter of hydrogen and $\frac{2}{3}$ liter of carbon dioxide; the 2-liter bottle will contain $\frac{2}{3}$ liter of hydrogen and $1\frac{1}{3}$ liters of carbon dioxide. In other words, the molecules of carbon dioxide so intermingle with the molecules of hydrogen that every part of each bottle is a perfect mixture containing 2 molecules of carbon dioxide to every molecule of hydrogen. The intermingling of gaseous molecules, without the aid of outside agencies, is called gaseous diffusion. The mixture resulting from the diffusion of gases

is a permanent one. Unlike liquids, the gases do not separate into layers in accordance with their densities.

43. Computing Density of Gas.—The quotient resulting from the weight of a given volume of a gas divided by the weight of an equal volume of some other gas, used as a standard, under the same conditions of temperature and pressure, is the density of that gas. Under standard conditions of temperature and pressure, 1 liter of oxygen weighs 1.429 g. Under the same conditions, 1 liter of air weighs 1.293 g. The density of oxygen, then, with reference to air as a standard, is $\frac{1.429}{1.293} = 1.105$. One liter of hydrogen at 0° C. and 760 mm. pressure weighs .08987 g. The density of oxygen with reference to hydrogen would then be $\frac{1.429}{.08987} = 15.912$. To state it differently, a given volume of oxygen weighs 1.105 times as much as a corresponding volume of air, and 15.912 times as much as an equivalent volume of hydrogen.

44. Graham's Law.—The speeds of diffusion vary with different gases. A detailed study of gaseous diffusion was made by the English chemist, Graham. From his work came the law which states that *the speeds of diffusion of gases are inversely proportional to the square roots of their densities.*

The relative densities of oxygen and hydrogen may be used as an example to demonstrate the law. It will be remembered that these densities are 15.912 and 1, respectively. The diffusion rate would then be $\frac{1}{\sqrt{15.912}}$ to $\frac{1}{\sqrt{1}}$ roughly, this is $\frac{1}{4}$ to 1. In other words, hydrogen diffuses about four times as fast as oxygen.

45. Kinetic-Molecular Theory.—The divisions of matter so far considered are based on a subdivision by physical and chemical means. There is also a classification based on the freedom with which molecules of matter move about and with their own energy attract one another. The power to attract one another,

or hold one another together, is called cohesion for like molecules, and adhesion for unlike molecules.

According to theory, each and every molecule has a vibratory motion. In a solid the molecules are held closely together and kept from getting away from one another by the attractive force among them. Still they are a considerable distance apart compared with their diameters.

In liquids, the distance through which a molecule may move is increased, and the attractive force that binds the molecules together is lessened.

In gases, the molecular vibrations have increased to such an extent that the distance between any two molecules has become too great for the attractive force to be effective. The molecules are following independent paths and are constantly colliding and rebounding in every direction.

46. Substances pass from one to another of these physical forms under changes of temperature and pressure. If, for example, a piece of ice is heated sufficiently, it melts and forms water. The molecules of ice, firmly held together by cohesion, can retain their original positions without being supported by some external rigid body; that is, a piece of ice will retain its shape without being supported by a container. The molecules of water, on the other hand, attract one another to a lesser extent than those of ice and are held together loosely; some external rigid container is needed to hold the liquid in position. Water held in a certain shape by a container will not keep that shape when poured into a container of another shape but will assume the shape of the second receptacle.

If water is heated further, it forms a gas known as steam. The molecules of steam tend to follow independent paths and fly off into space, unless held in a closed vessel.

The theory regarding the movement of molecules is known as the kinetic-molecular theory, the word *kinetic* signifying motion. According to the theory, the molecules of all matter, whether in the solid, liquid, or gaseous form, are in a state of constant motion. The kinetic-molecular theory literally means the theory of moving molecules.

47. With reference again to the case of a solid, the molecules are in such intimate contact with one another that great freedom of movement is prohibited. In other words, intermolecular spaces do not exist to any extent. This explains why solids retain their various shapes and why they possess such low compressibility, since, according to the theory, it is not the molecules in any substance that are compressed when the substance is subjected to great pressure, but the intermolecular spaces are lessened, causing a reduction in volume upon the application of pressure. But, if intermolecular spaces hardly exist, as in the case of a solid, its low compressibility is easily explained. In the case of a liquid, the intermolecular spaces are slightly larger than those of a solid and this means that the molecules of the liquid are more free to move; consequently, a liquid quickly takes the shape of its container. The spaces between the molecules of a gas at ordinary temperature and pressure must be relatively very great, compared with the spaces between the molecules of a solid or a liquid, and this accounts for the great compressibility of gases. The molecules of a gas are apparently in constant motion. They are traveling in straight lines in all directions, and are continually colliding with one another and with the walls of the containing vessel. Because of this bombardment of the walls of the containing vessel, a confined gas exerts a certain pressure. It is also assumed that, when the molecules of a gas collide, they rebound with the same velocity and energy as before. If this were not the case, they would soon lose their momentum and come to rest. Such a condition would result in a decrease in pressure of a confined gas, which does not occur. Therefore, it is assumed that the molecules of a gas are perpetually in motion.

LIQUEFACTION OF GASES

48. **Critical Temperature.**—It is reasonable to assume in the light of the knowledge gained so far that in the molecules of all gases there are two unseen forces that constantly oppose each other. The kinetic energy of the molecules is an inherent force that attempts to separate them one from the other. Opposing this force is the cohesive attraction tending to hold the

CRITICAL TEMP.

molecules together. The kinetic energy will vary with the temperature. Cohesive attraction appears to be relatively constant throughout temperature changes but varies considerably with different gases.

The minimum condition required to liquefy a gas is the equality of the cohesive force and the kinetic energy of the molecules. At high temperatures this equality can never take place, for the kinetic energy is too great and easily overcomes the cohesive attraction of any gas. However, as the temperature is reduced, the kinetic energy diminishes. A point is eventually reached where the two opposing forces are equal to each other and the gas is converted to the liquid state. The highest temperature at which a gas can be liquefied is known as the critical temperature of that gas. It should be apparent that critical temperatures vary with different gases.

49. Critical Pressure.—In order to liquefy a gas at the critical temperature, the surrounding conditions must be favorable at all times. Any slight change that would be favorable to the kinetic energy will prevent liquefaction of the gas. Conversely, any help that can be given the cohesive properties of the gas will hasten liquefaction. An agency that will bring the molecules closer together will directly aid the cohesive force. Pressure is the means of bringing about this condition. The pressure required to liquefy a gas at its critical temperature is known as the critical pressure.

50. Boiling Point.—Liquefaction of a gas is made easier as the temperature is lowered. This is easily understood, for, as the temperature is reduced, the kinetic energy diminishes, thus materially contributing to the forces of cohesion. If the temperature is continuously lowered, less and less pressure is required until finally the gas liquefies under ordinary atmospheric pressure. The temperature at which a gas is liquefied at atmospheric pressure is called the boiling point of that gas.

51. Liquid Air.—The manufacture of liquid air is an important industry, because it furnishes oxygen and nitrogen for many other industrial operations. The first operation in the manu-

facture of liquid air is the removal of carbon dioxide gas. By washing the air with water and then passing it over lime, the carbon dioxide is removed. The purified air is highly compressed at about 200 atmospheres. Most of the heat developed in compression is removed by running water. All the moisture that is present in the air is condensed at the high pressures encountered. The water plus the last traces of carbon dioxide are absorbed by sodium hydroxide. The apparatus for liquefying air is shown in Fig. 5.

52. The pump *a*, Fig. 5, draws air through the inlet pipe *m* into the apparatus. When a sufficient volume has been obtained, the inlet is closed by means of the valve shown at the junction

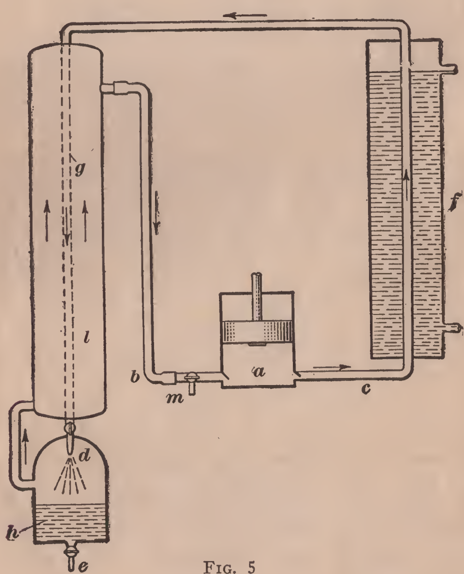


FIG. 5

of the pipes. The pump, which is able to exert a pressure of about 50 atmospheres, or about 735 pounds per square inch, forces the gas into the tube *c* in the direction of the arrows and it escapes, eventually, through the orifice *d* into the receiver *h*.

The heat developed by the compression of the gas in the pump is removed from the gas while passing through the tube *c*, the

first portion of this tube being cooled by the water-jacket *f*. The temperature of the gas is also lowered as a result of its expansion on leaving the orifice *d*, and a further reduction of temperature takes place as the gas expands on entering the large tank *l*. The cooled gas serves now as a means for cooling the gas descending through the tube *g*, which gas is again cooled by reason of its expansion into *h*. As the gas is continuously passing through the apparatus its temperature is gradually reduced, until finally the critical temperature is reached and the gas condenses. The liquid gas may be drawn off through the tube *e*.

PROPERTIES OF LIQUIDS

53. **Density.**—Different liquids vary greatly in density. The most common liquid in existence is water. It is used as the standard with which all other liquids are compared. Since a change in temperature affects the density of all liquids, a standard temperature must be selected at which their densities are determined.

At 4° C., water assumes its maximum density. The weight of 1 ml. of water at 4° C. is 1 gram. If it is warmed or cooled above or below this temperature, it gets lighter. Consequently, the density of water at 4° C. is selected as the standard and is designated as unity. The densities of all other liquids are compared with water at 4° C. For example, diethyl ether has a density of .708 at 25° C. In tables of physical constants it is usually shown as $.708\frac{25^\circ}{4^\circ}$. The term density is usually used synonymously with the expression specific gravity. Density is the weight per unit volume expressed in such terms as grams per milliliter or pounds per cubic foot. Specific gravity is defined as the ratio of the mass of a body to the mass of an equal volume of water at 4° C.

54. **Vapor Pressure and Boiling Point.**—The volatility, or rate of evaporation, is different with different liquids. If a small volume of ether is poured on a flat surface, it disappears rapidly, while an equivalent volume of water remains much longer. On the other hand, if sulfuric acid or glycerin is exposed

H_2SO_4

to the atmosphere it increases in volume owing to its tendency to absorb moisture.

In order to make clear the meaning of vapor pressure, it is assumed that a certain volume of liquid is placed in a tube and the tube closed and evacuated. Evaporation of the liquid starts immediately, in which the molecules fill the evacuated space above. Many of the gaseous molecules return to the liquid, the phenomenon being known as condensation. At a definite temperature the rates of evaporation and condensation become equal and a state of dynamic equilibrium exists. The pressure that is exerted by the vapor at this definite temperature is known as the vapor pressure.

When heat is applied to a liquid in a container, the vapor pressure increases. At a certain temperature the vapor pressure of the liquid becomes equal to the atmospheric pressure, and evaporation is visible. The liquid has reached the boiling point.

55. Surface Tension.—On the surface film of a liquid an unbalanced condition exists among molecules. This can be readily understood when one realizes that there is an attraction in all direction on the molecules within the liquid. If you watch water slowly drip from a faucet, you will observe that each drop becomes spherical at the moment it is detached. The formation of the spherical drops is due to surface tension. Surface tension is a force, acting in the surface of the liquid, that tends to diminish the surface or to prevent the surface from stretching. The smaller the drop of a liquid the less its tendency to be flattened by the pull of gravity and the more spherical will be the shape it approaches under the pull of surface tension.

PROPERTIES OF SOLIDS

56. Solidification of Vapors.—Almost all vapors, when cooled sufficiently under atmospheric pressure, will liquefy. However, there are a few that pass directly from the gaseous state to the solid state. Two notable examples of this type of change are ammonium chloride and arsenious oxide. There are other substances which exhibit this type of change, but only under certain conditions. For instance, water, sulfur, and iodine

may change from the gaseous state to either a liquid or a solid. Under certain conditions, atmospheric moisture may become rain or ice. Iodine, when kept in an enclosed space, will pass from the solid state to a gaseous state without passing through the liquid condition. This peculiar action is called sublimation. On the other hand, if iodine is heated slowly, it will form a liquid and then a gas. When cooled, the gas returns directly to the solid state without passing through the liquid state.

57. Solidification of Liquids.—At a temperature of 0° C. and a pressure of 1 atmosphere, water freezes and ice melts. When water is subjected to low temperatures, it steadily loses heat until 0° C. is reached. At this point there is no further change of temperature until all liquid water has frozen to ice. The temperature then continues downward. The same peculiarity exists when ice is melting. When ice starts to melt at 0° C. there is no further rise in temperature until all ice becomes liquid water. Heat breaks up the crystalline arrangement of the molecules of ice. The raising or lowering of the temperature of a gram of water through 1 degree involves the addition or removal of 1 calorie of heat. However, the conversion of 1 gram of water at 0° C. to a gram of ice at 0° C. requires the removal of 79 calories of heat. In order to convert 1 gram of ice at 0° C. to 1 gram of water at 0° C., an absorption of 79 calories of heat is required. This is known as the heat of fusion of ice.

VALENCE

58. Attraction of Atoms.—When observing the many varied chemical changes in nature, one is apt to wonder whether there is not some force or power that causes them. It is well to consider this subject before proceeding further with the study of chemical laws and definitions. When atoms of elements unite to form molecules of new compounds, they do so in such a manner that the forces of attraction of each atom for the other are satisfied. For instance, when the elements sodium and chlorine are brought together, the compound sodium chloride, *NaCl,* consisting of 1 atom of sodium and 1 atom of chlorine, is formed. The chemical affinity of each atom in this compound

is said to be satisfied, because it can be shown that the *Cl* will not unite with any additional atoms of *Na,* nor will the *Na* unite with additional atoms of *Cl.* In a similar manner, when calcium and chlorine are brought together, the compound calcium chloride, $CaCl_2$, is formed. Here 2 atoms of *Cl* are required to satisfy the chemical attraction, or affinity, which calcium has for chlorine.

59. What Is Valence?—The element hydrogen is frequently used in the study of chemistry as a standard of comparison. It was formerly used as a standard for the atomic weight system, the value assigned to it being 1. Hydrogen is also used as a standard of density for gases and, in the case of valence, hydrogen also serves as the standard and is given a value, or a valence, of 1, and the valences of all other elements are compared with it.

Valence should not be confused with chemical affinity, for each has its own distinctive chemical significance. Valence refers to the number of atoms of one element that can unite with an atom of another element, while chemical affinity treats of the relative force of attraction possessed by elements. An illustration will probably make this distinction clear. The elements chlorine and sodium have been found by experiment to have a greater chemical affinity for each other than have chlorine and calcium, yet 1 atom of calcium can unite with twice as many chlorine atoms as 1 atom of sodium, because calcium has a valence twice that of sodium. If all three elements are present, chlorine and sodium will unite much more readily than will chlorine and calcium, because the chemical affinity between chlorine and sodium is greater than that between chlorine and calcium. Therefore, it is obvious that chemical affinity and valence have entirely different meanings.

60. It can be shown that certain elements such as hydrogen, *H,* and chlorine, *Cl,* combine with each other, atom for atom, and form the compound hydrochloric acid, *HCl.* As previously stated, when atoms of different elements combine, the chemical affinity and the valence of each must be satisfied. Therefore, as the valence of *H* is taken as 1, that of chlorine must also be 1. The compounds sodium chloride, *NaCl,* and calcium chloride, $CaCl_2$, have also been mentioned. By applying the reasoning

used in establishing the valence of chlorine in the compound *HCl*, the conclusion is reached that, since chlorine has a valence of 1 in *HCl*, it may also have a valence of 1 in *NaCl*. This is actually the case, and the valence of *Na* must also be 1. In taking up the case of calcium in calcium chloride, it is noted that there are 2 atoms of chlorine, and, since each atom has a valence of 1, the total valence must be 2. Therefore, calcium has a valence of 2. In other words, calcium has a valence equivalent to 2 hydrogen atoms. $CuCl_2$

61. Valence is correctly defined as the smallest combining capacity of 1 atom of an element and is indicated by the number of hydrogen atoms with which an atom of the element in question will unite, or by the number of hydrogen atoms, or its equivalent, that an atom of the element will replace.

The following illustrations will make the definition clear: The compound water has the formula H_2O. The valence of hydrogen is 1 and, as there are 2 atoms of hydrogen in this formula, there is also a total of 2 valences for hydrogen. Since the valences of each element must be satisfied, the valence of oxygen must also be 2, and it is worth noting here that, in all its compounds, oxygen has always been found to have a valence of 2. As another example, use the compound sodium oxide, the formula for which is Na_2O. We have already established the valence of sodium in the compound sodium chloride, *NaCl*, as being 1 and that of oxygen in H_2O as being 2. Therefore, when sodium and oxygen are brought together and react, a compound of sodium and oxygen is formed. As oxygen has a valence of 2, two atoms of sodium with a total valence of 2 must be present in the compound.

62. Inspection of Table VII shows that some of the elements have more than one valence. In other words, it will be found that the valence of some of the elements listed depends on the elements with which they combine. A variable valence should not be confusing if the valences of a few of the common elements are remembered. Thus, hydrogen, *H*, always has a a valence of 1; oxygen, *O*, a valence of 2; and sodium and potassium, a valence of 1 each.

TABLE VII

USUAL VALENCE OF ELEMENTS

Name of Element	Symbol	Valence	Name of Element	Symbol	Valence
Aluminum	*Al*	3	Nickel	*Ni*	2, 3
Antimony	*Sb*	3, 5	Nitrogen	*N*	1, 2, 3, 4, 5
Arsenic	*As*	3, 5	Osmium	*Os*	2, 3, 4, 6, 8
Barium	*Ba*	2	Oxygen	*O*	2
Beryllium	*Be*	2	Palladium	*Pd*	2, 3, 4
Bismuth	*Bi*	3, 5	Phosphorus	*P*	3, 5
Boron	*B*	3	Platinum	*Pt*	2, 4
Bromine	*Br*	1, 3, 5, 7	Potassium	*K*	1
Cadmium	*Cd*	2	Praseodymium	*Pr*	3
Calcium	*Ca*	2	Rhodium	*Rh*	2, 3
Carbon	*C*	2, 3, 4	Rubidium	*Rb*	1
Cerium	*Ce*	3, 4	Ruthenium	*Ru*	3, 4, 6, 8
Cesium	*Cs*	1	Samarium	*Sm*	3
Chlorine	*Cl*	1, 3, 5, 7	Scandium	*Sc*	3
Chromium	*Cr*	2, 3, 6	Selenium	*Se*	2, 4, 6
Cobalt	*Co*	2, 3	Silicon	*Si*	4
Copper	*Cu*	1, 2	Silver	*Ag*	1
Erbium	*Er*	3	Sodium	*Na*	1
Fluorine	*F*	1	Strontium	*Sr*	2
Gallium	*Ga*	2, 3	Sulfur	*S*	2, 4, 6
Germanium	*Ge*	2, 4	Tantalum	*Ta*	3, 4, 5
Gold	*Au*	1, 3	Tellurium	*Te*	2, 4, 6
Hydrogen	*H*	1	Terbium	*Tb*	3
Indium	*In*	3	Thallium	*Tl*	1, 3
Iodine	*I*	1, 3, 5, 7	Thorium	*Th*	4
Iridium	*Ir*	3, 4	Tin	*Sn*	2, 4
Iron	*Fe*	2, 3	Titanium	*Ti*	3, 4
Lanthanum	*La*	3	Tungsten	*W*	2, 4, 6
Lead	*Pb*	2, 4	Uranium	*U*	4, 6
Lithium	*Li*	1	Vanadium	*V*	2, 3, 4, 5
Magnesium	*Mg*	2	Ytterbium	*Yb*	3
Manganese	*Mn*	2, 3, 4, 6, 7	Yttrium	*Y*	3
Mercury	*Hg*	1, 2	Zinc	*Zn*	2
Molybdenum	*Mo*	3, 4, 5, 6	Zirconium	*Zr*	4
Neodymium	*Nd*	3			

Among the common elements which have a variable valence are chlorine, Cl, and nitrogen, N. The particular valence of either of these elements in any of their compounds can easily be worked out. Some of the more common compounds which nitrogen forms are nitrous oxide, N_2O; nitric oxide, NO; nitrogen dioxide, NO_2; nitrogen trioxide, N_2O_3; and nitrogen pentoxide, N_2O_5. In N_2O nitrogen has a valence of 1, as oxygen always has a valence of 2, and $2N$ with a total valence of 2 are required to satisfy the valence of oxygen. In NO the valence of N must be 2, since 1 atom of it has combined with 1 atom of oxygen, which has a valence of 2. In NO_2 the valence of N is 4, since 2 atoms of oxygen have a total valence of 4. In N_2O_3 there are 6 valences for 3 oxygen atoms, and since there are only $2N$, the valence of each must be 3. In N_2O_5 there is a total of 10 valences for oxygen. Therefore each N must have a valence of 5.

Nitrogen is also a constituent of nitric acid, the formula of which is HNO_3. Here there are 6 valences for oxygen and 1 for hydrogen, leaving 5 for nitrogen.

TABLE VIII

DIFFERENT VALENCES AND THEIR DESIGNATIONS

Valence	Element is said to be	Element is called a
1	Monovalent, or univalent	Monad
2	Divalent, or bivalent	Dyad
3	Trivalent, or tervalent	Triad
4	Tetravalent, or quadrivalent	Tetrad
5	Pentavalent, or quinquivalent	Pentad
6	Hexavalent, or sexivalent	Hexad
7	Heptavalent, or septivalent	Heptad
8	Octavalent	Octad

The compounds of chlorine in which the valence varies are also numerous, but the valence can be worked out in exactly the same manner as for nitrogen. Thus, in the compound potassium chlorate, $KClO_3$, there are 6 valences for 3 oxygen atoms, 1 valence for potassium, leaving 5 for chlorine. In the

compound sodium perchlorate, $NaClO_4$, there are 8 valences for oxygen, 1 for sodium, leaving 7 for chlorine.

63. **Designation of Valence.**—The subject of variable valence has been sufficiently explained to make it clearly understood, and it is now merely necessary to show by means of Table VIII how the valence of an element is designated.

CHEMICAL EQUATIONS

CONSTRUCTION

64. **Chemical Reactions.**—When the subject of chemical and physical changes was discussed, it was stated that, when a substance has been changed to an altogether different substance having a chemical composition entirely different from the original substance, a chemical change has taken place. We can now give this so-called chemical change the name by which it is always designated. In other words, when the chemical composition of a substance has been changed, a chemical reaction has taken place.

65. Up to this point a single substance only has been dealt with, but chemical reactions are not limited to single substances. In fact, the majority of reactions occur between two or more substances. When two substances are heated together and a chemical reaction occurs, the identity of each substance is lost and a new product, called the product of the reaction, is formed. For example, if a quantity of iron filings is mixed with powdered sulfur or flowers of sulfur and the mixture is strongly heated, a compound of iron and sulfur is formed. This compound, known as ferrous sulfide, would have properties differing greatly from those of the original iron or sulfur.

It has already been shown that elements and compounds are represented by symbols and formulas. In the writing of chemical equations, these are employed.

66. In the case of the reaction between iron and sulfur by which ferrous sulfide is formed, the statement might merely be made that, when iron and sulfur are heated together, the product

is ferrous sulfide, but by representing this reaction by means of the following equation:

$$Fe + S + \text{heat} \longrightarrow FeS$$

several advantages are gained. At a glance this equation shows to a chemist that, when iron and sulfur are heated together, ferrous sulfide is formed. Ordinarily, the word *heat* would not be written into the equation but would be symbolized by $\triangle$. Formerly it was the custom to show the reaction by means of the equality sign, hence the term *equation*. The arrow has replaced the equality sign and has the same meaning. It should be read "forms."

67. The question is often asked, "How are the products of reaction identified, or known?" This question can be answered only by saying that in the beginning of the science of chemistry it was not known what was formed when different substances reacted. Therefore, it was necessary to devise methods of analysis so that these new products might be identified. In this way, the products formed as the result of most chemical reactions are known, and, if they are not known, their identity and composition must be established by methods peculiar to analytical chemistry.

68. Examination of the equation in Art. 66 shows that the number of iron and sulfur atoms on each side of the arrow is the same. In other words, the equation may be said to balance. This is an important point to remember, since chemical equations cannot be correct unless they balance.

69. Facts Concerning Chemical Equations.—Before going deeper into the subject of equation writing, it will be well to introduce a few rules for the application and guidance of the beginner in chemistry.

1. A true chemical equation is a statement of a fact. Therefore, if an equation is written for a reaction that cannot take place, the equation is incorrect. In other words, chemical equations are correct only when the changes they represent actually take place.

2. The number of atoms of any element that appears on each side of the equality sign of an equation must be the same.

3. The valence of the elements must be observed when the symbols representing the elements are placed in an equation.

4. The products of most reactions have been analyzed and determined, and must be remembered.

Apply the four points just given to the only equation which thus far has been studied, namely: $Fe+S \xrightarrow{\triangle} FeS$.

In the first place, it is known that the reaction takes place, so the equation may be written to represent it. With respect to the second point, the number of atoms of each element that appears on each side of the equality sign is the same. In the third place, the symbols for iron and sulfur are written in accordance with what is known concerning the valence of each. It is also known that FeS is formed, because the product has been analyzed. Therefore, it is seen that the equation is correctly written.

70. Balancing Simple Equations.—The equations here chosen to illustrate how simple chemical equations may be balanced conform to two of the statements previously made; that is, they represent reactions which actually take place and the products of each reaction are known.

The first equation to be considered is that representing the reaction which takes place between sodium chloride, $NaCl$, and sulfuric acid, H_2SO_4. The products of this reaction are sodium sulfate, Na_2SO_4, and hydrochloric acid, HCl. The incomplete, or skeleton, equation may be written:

$$NaCl+H_2SO_4 \longrightarrow Na_2SO_4+HCl$$

It will readily be seen that this equation does not balance for the following reasons: On the left-hand side of the equation there is only 1 sodium (Na) atom while on the right-hand side there are 2 sodium (Na) atoms. Again, on the left there are 2 hydrogen (H) atoms and on the right there is only 1 hydrogen (H) atom. To complete this equation it is necessary to cause the Na atoms to balance by writing $2NaCl$ on the left. Thus:

$$2NaCl+H_2SO_4 \longrightarrow Na_2SO_4+\mathrm{HCl}$$

As the result of this addition there are now 2 chlorine (Cl) atoms on the left and only 1 chlorine (Cl) atom on the right, and it is necessary to balance both the H and the Cl atoms. This is easily accomplished by writing $2HCl$ on the right-hand side of the equation. Thus:

$$2NaCl+H_2SO_4 \longrightarrow Na_2SO_4+2HCl$$

In the equation now, the number of atoms of each element on each side of the arrow is the same. In other words, the equation is balanced.

71. As a second example, consider the formation of hydrogen and ferrous chloride, $FeCl_2$, when hydrochloric acid is brought in contact with iron, Fe. The skeleton equation may be written as follows:

$$Fe+HCl \longrightarrow FeCl_2+H_2$$

In this case there are 2 chlorine atoms and 2 hydrogen atoms on the right and only 1 of each on the left, but to cause the equation to balance it is necessary to write $2HCl$ on the left-hand side. Thus:

$$Fe+2HCl \longrightarrow FeCl_2+H_2$$

In this equation all of the atoms balance.

72. Use of Valence in Writing Chemical Equations.—If the principles of valence have been thoroughly studied, their application to the writing of correct chemical equations should present little or no difficulty. It is not necessary, in balancing some types of equations, to consider the principles of valence. Other types, however, cannot be balanced without considering these principles. Therefore, the application of valence to the writing of simple equations will be illustrated at this point in order to render the work of balancing more complicated equations less difficult. Consider the reaction which takes place when iron rusts. In order to make the work as simple as possible, the rusting of iron may be said to be due to the action of atmospheric oxygen on the iron, the product of the reaction being ferric oxide, Fe_2O_3. The skeleton equation for this reaction is written as follows:

$$Fe+O \longrightarrow Fe_2O_3$$

Oxygen, however, is diatomic and is written O_2, so the equation becomes

$$Fe + O_2 \longrightarrow Fe_2O_3$$

This equation can be balanced by the methods given in preceding examples, but not so readily, and the processes previously used applied to this equation resolve themselves into hit-or-miss methods, which, as far as practical results are concerned, are useless.

For example, an attempt might be made to balance the equation by first making the Fe atoms equal. Thus:

$$2Fe + O_2 \longrightarrow Fe_2O_3$$

It is not so simple a matter to make the oxygen atoms balance. In fact, it can only be done by using a fractional coefficient, $1\frac{1}{2}$, before the O_2, which would give 3 oxygen atoms on the left. Fractional coefficients are not used in equation writing. When the equation has been written

$$2Fe + 1\tfrac{1}{2}O_2 \longrightarrow Fe_2O_3$$

it will be necessary to multiply it throughout by 2 in order to eliminate the fractional coefficient. When this has been done the equation will be correctly balanced and should appear as

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

When an equation cannot be readily balanced by other methods, the process becomes a short and simple one by using the valences of the elements. Thus, in the equation

$$Fe + O_2 \longrightarrow Fe_2O_3$$

Fe on the left has a valence of zero, since it is not in combination, while on the right, Fe has a valence of 3, or there is a total difference of 3 in valence. This difference must have been due to the action of oxygen, so the 3 is placed before the O_2 and the equation at once becomes

$$Fe + 3O_2 \longrightarrow Fe_2O_3$$

Now, the preceding methods may be used, and to have 6 oxygen atoms on the right, Fe_2O_3 must be multiplied by 2, which

gives $4Fe$ atoms on the right. It is now necessary that $4Fe$ also appear on the left. Thus:

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

73. In connection with the subject of equation writing, a question frequently asked by those beginning the study of the subject of chemistry is, "How can I tell what the products of a certain reaction will be?" The answer to this question is that where one's chemical experience is very limited, the memory must be depended on largely or reference made to the text to determine the course of a certain reaction. As one gradually becomes familiar with all types of chemical reactions, a certain similarity will be noticed among them. For instance, it will be seen that when acids act on carbonates the usual products are in part carbon dioxide and water. Thus:

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$$

The temperature at which a reaction takes place, whether in a hot or a cold solution, sometimes determines whether certain products or others will be formed; and again, quantities of materials, strengths of acids and solutions, and other factors determine what will happen when compounds react.

74. The ability to predict what will be formed as the result of a simple chemical reaction depends on one's knowledge of the effect of the foregoing factors. This knowledge can be gained only by a thorough study of all of the reactions represented by equations given in the text. If this procedure is followed, a beginner will soon find that he has learned all of the principal reactions, and eventually all of those which he meets in subsequent work will in most respects be similar to those which he has already studied.

TYPES OF CHEMICAL REACTION

75. **Direct Combination.**—In the type of reaction known as direct combination, two or more substances are first used to form a compound. Of necessity, man must keep warm in the cold seasons and must cook his food in any season. Both are accomplished in most cases by the burning of coal. The burning

$C + O_2$

of coal is essentially a rapid combination of carbon and oxygen to form a gas known as carbon dioxide. When lime is slaked, another rapid combination is effected in which heat is produced. In the latter instance, the compound calcium oxide combines with the compound water to form the compound calcium hydroxide. The two materials simply unite and form one compound with nothing left over. The foregoing examples portray the simplest type of chemical reaction and are called chemical union, or combination. Chemical union is the combination of simple substances to form a more complex one.

CO_2

$+ H_2O$

2. **76. Simple Decomposition.**—Simple decomposition involves only one substance, which breaks down by a chemical reaction into two or more simpler substances. The decomposition of salt ($NaCl$) into sodium and chlorine is an example of this type of reaction. The decomposition of potassium chlorate and of mercuric oxide by heat are also examples. Thus:

$$2KClO_3 \xrightarrow{\Delta} 2KCl+3O_2$$

$$2HgO \xrightarrow{\Delta} 2Hg+O_2$$

3- **77. Simple Displacement.**—Sulfuric acid is a compound containing three elements: hydrogen, sulfur, and oxygen. If the metal zinc is added to the acid, gas bubbles appear throughout the solution. The zinc disappears in time and the bubbles are found to be the inflammable gas, hydrogen. The white solid that forms is found to contain zinc, sulfur, and oxygen. It is seen at once that the zinc has taken the place of the hydrogen to form a compound known as zinc sulfate. This is called a displacement reaction. It is a reaction in which one element displaces another in a compound to set the latter element free.

78. Double Displacement or Decomposition.—The most common type of chemical reaction is that in which two compounds interact and produce two other compounds. This is called double decomposition, because each compound apparently breaks up into two parts, each of which unites with a different part of the other compound. For example, take the reaction

of sulfuric acid, H_2SO_4, on marble, $CaCO_3$, to form calcium sulfate, $CaSO_4$, and carbonic acid, H_2CO_3.

$$H_2SO_4 + CaCO_3 \longrightarrow CaSO_4 + H_2CO_3$$

In this type of reaction the first part of the first compound unites with the second part of the second compound, and the second part of the first compound unites with the first part of the second compound.

79. Chemical Rearrangement.—Infrequently a fifth type of reaction occurs. A few compounds exist that exhibit the ability to rearrange their constituent elements to form entirely new and different compounds under different conditions. There is no gain or loss of material in the process. Ammonium cyanate is a compound typical of this type of reaction. When a water solution of the material is allowed to evaporate under ordinary conditions, the residue is entirely different from the starting compound. The constituents of ammonium cyanate, NH_4CNO, have completely rearranged themselves to form urea, NH_2CONH_2. A chemical change in which the resulting product contains the same constituents in the same weight proportions as the original substance, but in different relationship one with the other, is called rearrangement.

NOMENCLATURE

ACIDS, BASES, AND SALTS

80. General Remarks.—In the early days of the development of chemical science, names were applied to substances without following any fixed plan. Now, substances referred to in chemistry are named according to a plan in which the chemical composition of the material plays the most important part. No one can consider his education in chemistry satisfactory unless he thoroughly understands the system followed in naming substances and knows how to apply it. One should know the proper use of names like acid, base, salt, peroxide, sulfide, chlorate, chlorite, and many other terms. The following explanation of this subject should be considered carefully, for it treats of the

definitions and rules upon which the naming of substances is based and is applicable to every branch of inorganic chemistry.

81. Acids.—It is well to consider some of the acids to determine whether there are any specific characteristics that distinguish them from all other substances. Sulfuric acid and hydrochloric acid are two substances used a great deal by chemists, and their formulas, H_2SO_4 and HCl, respectively, show that they both contain hydrogen; but so do many substances that are not acids, as, for example, calcium hydroxide, $Ca(OH)_2$, potassium hydroxide, KOH, and water, H_2O. How does the hydrogen in acids differ from the hydrogen in other substances?

Acids, when allowed to act on metals, as, for example, zinc or iron, chemically combine with them, and hydrogen is liberated. The following reactions take place when zince reacts with each acid mentioned:

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$
$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

In the first case, zinc chloride and hydrogen are formed; in the second, zinc sulfate and hydrogen are the products of the reaction. It has been found by experiment that acids are the only substances of those containing hydrogen that can react with metals or equivalent elements to form the products hydrogen and salts. Therefore, acids contain hydrogen that can be replaced by metals.

It has also been found that acids can change the color of litmus, a vegetable dye of indefinite chemical composition, from blue to red.

From the foregoing facts, it is obvious that acids have distinct characteristics possessed by no other class of substances, and that the following definition will apply: An acid is any substance containing hydrogen that is replaceable by a metal and that will also change the color of litmus from blue to red.

82. Bases.—A base is a substance that imparts a soapy feeling when touched, turns red litmus blue, and reacts with an acid to form a salt and water. A base may be defined as a

substance that, when dissolved in water, dissociates to form an electropositive metallic atom or group and the electronegative hydroxyl group. The hydroxyl group consists of 1 atom of oxygen and 1 atom of hydrogen and is often written (OH^-). It can exist as such only in water. It is the hydroxyl group that imparts the characteristic properties of all bases. Bases that are freely soluble are called alkalies. Some of the well-known alkalies are sodium hydroxide, $NaOH$, potassium hydroxide, KOH, ammonium hydroxide, NH_4OH, and barium hydroxide $Ba(OH)_2$.

83. **Metals and Non-Metals.**—Elements that have a characteristic luster, conduct electricity, are more or less malleable, and form stable salts with acids are called metals. They are further distinguished by forming basic oxides, or oxides that are capable of reacting with water to form bases. Calcium oxide is typical of a basic oxide. When it reacts with water it forms the base, calcium hydroxide. Furthermore, an element that reacts with other elements or substances to form a compound in which the element appears as the positive part of the molecule is called a metal.

Non-metals do not have a metallic luster. Their oxides and hydroxides generally give acid reactions. In combination with hydrogen or hydrogen and oxygen, they all form acids.

The dividing line between metals and non-metals is not a sharp one. Some elements can act either as metals or non-metals, depending on the conditions. For example, silicon appears as a metal in silicon disulfide, SiS_2, and as a non-metal in metasilicic acid, H_2SiO_3. Elements that commonly act both as metals and non-metals are called metalloids. Metals are called positive elements and non-metals negative elements, based on their action when an electric current is passed through a solution containing them. The procedure is called electrolysis.

Table IX shows the divisions into which the important elements are conveniently grouped. It should be borne in mind that the classification is somewhat arbitrary, for some elements classed as metals will, under certain conditions, act as non-metals. For example, aluminum ordinarily acts as a metal;

TABLE IX
METALS AND NON-METALS

Metals		Non-Metals
Names	Names	Names
Aluminum	Neodymium	Arsenic
Antimony	Nickel	Boron
Beryllium	Osmium	Bromine
Bismuth	Palladium	Carbon
Cadmium	Platinum	Chlorine
Calcium	Potassium	Fluorine
Cerium	Praseodymium	Iodine
Cesium	Radium	Nitrogen
Chromium	Rhodium	Oxygen
Cobalt	Rubidium	Phosphorus
Columbium	Ruthenium	Selenium
Copper	Scandium	Silicon
Dysprosium	Sodium	Sulfur
Erbium	Silver	Tellurium
Europium	Strontium	
Gallium	Tantalum	
Gadolinium	Terbium	
Germanium	Thallium	
Gold	Thorium	
Holmium	Thulium	
Hydrogen	Tin	
Indium	Titanium	
Iridium	Tungsten	
Iron	Uranium	
Lanthanum	Vanadium	
Lead	Ytterbium	
Lithium	Yttrium	
Magnesium	Zinc	
Manganese	Zirconium	
Mercury		
Molybdenum		

yet, when its compounds react chemically with sodium hydroxide, they form sodium aluminate, in which compound aluminum acts as a non-metal, thus:

$$Al(OH)_3 + NaOH \longrightarrow NaAlO_2 + 2H_2O$$

84. **Neutralization.**—A neutral solution of the vegetable dye, litmus, is blue in color. As previously stated, solutions of acids have the power of changing this color to red and solutions of bases in water have the power of restoring the blue color. For example, if 2 or 3 drops of hydrochloric acid are added to a neutral litmus solution, the color of the solution will be changed to a bright red. If, now, a few drops of a solution of sodium hydroxide in water are added to the reddened litmus solution, the original blue color will be restored. By adding sufficient acid, the color of the solution may again be changed to red. When the red color is obtained, the solution is said to show an acid reaction, and when the blue color is restored, it shows an alkaline reaction. If the acid is very carefully added to the blue solution, the blue color will become less distinct at a certain point, and the solution, while still blue, will be found to have a slight reddish tint. At about this point the solution is neutral; that is, it is neither basic nor acid, the base and acid having neutralized, or destroyed, the properties of each other. This process of destroying basic and acid properties by allowing a base and acid to act on each other is known as neutralization.

The question now arises: Is a definite amount of acid required to neutralize a fixed amount of base? That the answer is affirmative may easily be shown by experiments and it is upon this principle that one of the branches of analytical chemistry is based.

85. As an example, if two solutions are prepared, one containing sulfuric acid, H_2SO_4, and the other sodium hydroxide, $NaOH$, and a volume of the acid solution is added to a volume of the alkaline solution containing a few drops of a solution of litmus, until the blue color of the litmus just changes to a red, the solution will be about neutral. The volumes of alkali and acid solutions used are then noted and a second experiment is performed, using the solutions in the same proportion but in different volumes, and it will be found, in each case, that a certain volume of the acid will neutralize a certain volume of the alkali.

86. **Salts.**—The term *salt* is ordinarily applied to sodium chloride. However, in chemistry it refers to the products

formed when the hydrogen atoms of acids have been replaced by metals. The majority of substances included in inorganic chemistry come under this heading. When zinc reacts with sulfuric acid, hydrogen and the salt, zinc sulfate, are formed, as here shown:

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

The metal portion of a salt, in this case zinc, is called the positive radical, and the acid portion with which the metal is combined, in this case SO_4, is called the negative radical. The negative radical of an acid is always the portion of the acid that combines with hydrogen. For example, the negative radical of nitric acid, HNO_3, is NO_3. A negative radical must be combined with a positive radical before its valence is satisfied.

Salts are commonly prepared by treating a base with an acid, and as a result of the reaction which takes place water is also produced. For example, potassium hydroxide and hydrochloric acid react to form the salt, potassium chloride, and water. Thus:

$$KOH + HCl \longrightarrow KCl + H_2O$$

Potassium chloride is found to have neither the basic properties of potassium hydroxide nor the acid properties of hydrochloric acid; in fact, it does not affect litmus in any way. It is a neutral substance so far as acid or basic properties are concerned.

87. Normal, Acid, Basic, and Double Salts.—The salts so far considered are those in which just enough metal is supplied to replace the hydrogen atoms in an acid; but there are other kinds. The following distinctions must, therefore, be made: A normal salt is a salt obtained from an acid by replacing all its hydrogen atoms by metal atoms. An acid salt is a salt obtained from an acid by replacing part of the hydrogen atoms by metal atoms. A basic salt is a salt obtained by the partial neutralization of a base by an acid. Following are equations showing the processes as used to produce the different kinds of salts:

A normal salt:

$$2KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$$

The resulting potassium sulfate, K_2SO_4, is a normal salt, for all the hydrogen atoms in the acid have been replaced by potassium atoms.

An acid salt:

$$KOH + H_2SO_4 \longrightarrow KHSO_4 + H_2O$$

The resulting potassium hydrogen sulfate, $KHSO_4$, is an acid salt, for not all the hydrogen atoms in the acid have been replaced by potassium atoms.

A basic salt:

$$Zn(OH)_2 + HCl \longrightarrow Zn(OH)Cl + H_2O$$

The resulting zinc hydroxy-chloride is a basic salt, for the base has been only partly neutralized by the acid.

88. A double salt differs from a normal salt in that the hydrogen atoms in the acid have been replaced by two or more metals. For example, the mixed salt sodium potassium sulfate, $KNaSO_4$, can be made by first producing the acid salt potassium hydrogen sulfate and then treating this with sodium hydroxide. Thus:

$$KOH + H_2SO_4 \longrightarrow KHSO_4 + H_2O$$
$$KHSO_4 + NaOH \longrightarrow KNaSO_4 + H_2O$$

A double salt is formed by the replacement of the hydrogen atoms in an acid by different metals.

ACID SALTS, BASIC ACIDS, AND ACID BASES

89. Nomenclature of Acid Salts.—Different names are applied to acid salts in accordance with the number of hydrogen atoms that have been replaced by metals. For example, phosphoric acid, H_3PO_4, can form three kinds of salts with sodium, depending on the number of sodium atoms that enter into the compound. Following are reactions showing the formation of the three salts formed by the action of phosphoric acid on sodium hydroxide:

$$NaOH + H_3PO_4 \longrightarrow NaH_2PO_4 + H_2O$$
$$2NaOH + H_3PO_4 \longrightarrow Na_2HPO_4 + 2H_2O$$
$$3NaOH + H_3PO_4 \longrightarrow Na_3PO_4 + 3H_2O$$

The salt having the formula Na_3PO_4 is called normal sodium phosphate, because all of the hydrogen atoms in the acid have been replaced by sodium atoms. The other two salts formed, Na_2HPO_4 and NaH_2PO_4, are acids salts, because only part of the hydrogen atoms in the acid have been replaced by sodium atoms. Different names must be applied to them so that they can be distinguished from each other.

90. Based on the number of sodium atoms in the various salts, the following system of naming them is used:

NaH_2PO_4 = monosodium phosphate
Na_2HPO_4 = disodium phosphate
Na_3PO_4 = trisodium phosphate

Based on the number of hydrogen atoms in the various salts, the following system of names is used:

NaH_2PO_4 = dihydrogen sodium phosphate
Na_2HPO_4 = monohydrogen sodium phosphate
Na_3PO_4 = normal sodium phosphate

Following is another set of names that is sometimes used:

NaH_2PO_4 = primary sodium phosphate
Na_2HPO_4 = secondary sodium phosphate
Na_3PO_4 = tertiary sodium phosphate

The name sodium phosphate is usually applied to Na_2HPO_4, for this salt is the most common of the three.

91. Some acids, like sulfuric acid, form but one acid salt with a metal. For example, potassium hydroxide can react with sulfuric acid in the following ways:

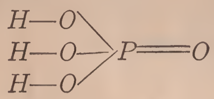
$$KOH + H_2SO_4 \longrightarrow KHSO_4 + H_2O$$
$$2KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$$

The normal salt, K_2SO_4, is called potassium sulfate. The acid salt is called acid potassium sulfate or potassium hydrogen sulfate. No other names are necessary, as in the case of the sodium phosphates, for there is but one acid potassium sulfate. Acid potassium sulfate is also called potassium bisulfate to show that twice as much SO_4 radical is present as is needed to form a normal salt with the potassium contained.

92. Basicity of Acids.—All the hydrogen atoms in the acids so far considered can be replaced by metals. There are a few acids in which only a part of the hydrogen atoms can be replaced, and, in order to explain more clearly why only certain of the hydrogen atoms are replaced, graphic or structural formulas, in which the valence of the different atoms is indicated by means of dashes, will be used.

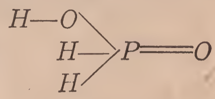
Consider the two oxyacids (acids containing oxygen), phosphoric acid, H_3PO_4, and hypophosphorous acid, H_3PO_2. Investigation in the laboratory has shown that, although phosphoric and hypophosphorous acids each contain 3 hydrogen atoms, only 1 of these in hypophosphorous acid may be replaced by a metal. The graphic or structural formulas of these acids are, respectively:

phosphoric acid



and

hypophosphorous acid



In these formulas the valence of each atom is indicated as follows: Hydrogen is always monovalent in any compound, therefore each hydrogen atom is shown, connected by one bond, or dash, to an oxygen or phosphorus atom. Oxygen is always bivalent and each atom is therefore shown connected once to hydrogen and once to phosphorus, thus satisfying its valence of 2. In each formula, however, one oxygen atom must be written with a double bond in order to show its valence of 2. Phosphorus in phosphoric acid is pentavalent and is shown connected twice to one oxygen and once to each of three oxygens, thus satisfying its valence of 5. The 3 hydrogen atoms of phosphoric acid are replaceable; the graphic formula is therefore written to show that each hydrogen atom is directly connected to oxygen, but only 1 of the hydrogen atoms of hypophosphorous acid is replaceable. The first acid contains three hydroxyl, *OH*, groups and the latter but one. Investigations showed that only hydrogen atoms in the hydroxyl groups of oxyacids are replaceable by metals, and that those hydrogen atoms of oxyacids that are not in the hydroxyl groups are not

replaceable by metals. These facts do not, of course, apply to acids that do not contain oxygen. The hydrogen atoms in acids like hydrochloric acid, HCl, and hydrofluoric acid, HF, can be entirely replaced by a metal.

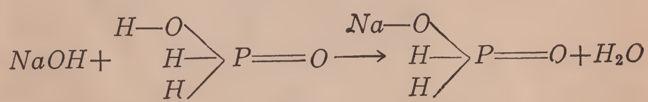
93. An acid that requires 1 atom of a monovalent metal to replace all the replaceable hydrogen in 1 molecule of it is said to be a monobasic acid. An acid that requires 2 atoms of a monovalent metal or 1 atom of a divalent metal to replace all the replaceable hydrogen in 1 molecule of it is said to be dibasic. An acid that requires 3 atoms of a monovalent metal or 1 atom of a trivalent metal to replace all of the replaceable hydrogen in 1 molecule of it is said to be tribasic. The basicity of an oxyacid, as shown before, is limited by the number of hydroxyl groups it contains; all the hydrogen atoms in acids that do not contain oxygen are replaceable by metals. The term *basicity* is used because the metals that replace the hydrogen in acids are usually obtained from bases.

94. Examples of Reactions Between Bases and Acids. The following examples will serve to make the conception of the basicity of acids clear:

Monobasic acids:

(*a*) $$NaOH + HCl \longrightarrow NaCl + H_2O$$

Hydrochloric acid is monobasic because 1 molecule of the base, sodium hydroxide, is needed to supply the 1 atom of monovalent sodium required to replace the hydrogen atom in one of its molecules.

(*b*) 

Hypophosphorous acid is monobasic because 1 molecule of the base, sodium hydroxide, is needed to supply the 1 atom of monovalent sodium required to replace the replaceable hydrogen atom in 1 molecule of it.

Dibasic acids:

(*a*) $$2KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$$

Sulfuric acid is dibasic because 2 molecules of the base, potassium hydroxide, are needed to supply the 2 atoms of monovalent potassium required to replace the replaceable atoms of hydrogen in 1 molecule of it.

(*b*) $$Ca(OH)_2+H_2SO_4 \longrightarrow CaSO_4+2H_2O$$

Sulfuric acid is shown to be dibasic by this reaction also, because 1 molecule of the base, calcium hydroxide, is needed to supply the 1 atom of divalent calcium required to replace the replaceable hydrogen atoms in 1 molecule of it.

Tribasic acids:

(*a*) $$3NaOH+H_3PO_4 \longrightarrow Na_3PO_4+3H_2O$$

Phosphoric acid is tribasic because 3 molecules of the base, sodium hydroxide, are needed to supply the 3 atoms of univalent sodium required to replace the replaceable hydrogen atoms in 1 molecule of it.

(*b*) $$Al(OH)_3+H_3PO_4 \longrightarrow AlPO_4+3H_2O$$

Phosphoric acid is shown to be tribasic by this reaction also, because 1 molecule of the base, aluminum hydroxide, is needed to supply the 1 atom of trivalent aluminum required to replace the replaceable hydrogen atoms in 1 molecule of it.

95. Acidity of Bases.—Bases are classed according to the number of hydrogen atoms that can be replaced in an acid by 1 atom of the metal supplied. For example, sodium hydroxide, $NaOH$, is a monacid base, because each molecule of it supplies 1 atom of monovalent sodium, capable of replacing 1 hydrogen atom in an acid. Calcium hydroxide, $Ca(OH)_2$, is a diacid base, because each molecule of it supplies 1 atom of divalent calcium, capable of replacing 2 hydrogen atoms in an acid. Aluminum hydroxide, $Al(OH)_3$, is a triacid base, because each molecule of it supplies 1 atom of trivalent aluminum, capable of replacing 3 hydrogen atoms in an acid. The reactions used to exemplify the basicity of acids should be studied in connection with the acidity of bases, for they show how monacid, diacid, and triacid bases react with acids.

COMPOUNDS CONTAINING TWO ELEMENTS

96. Binary Compounds.—Compounds that consist of two elements, or two groups acting as elements, are called binary compounds. The ending ide is applied to these substances. In Table X are shown a few examples of binary compounds and the system of naming them. In each case, the ending *ide* is applied to the negative elements; usually, a non-metal. The more positive element, usually a metal, is written before the negative element.

TABLE X
NOMENCLATURE OF BINARY COMPOUNDS

Positive Element	Negative Element	Formula	Name of Binary Compound
Magnesium	Oxygen	MgO	Magnesium oxide
Zinc	Sulfur	ZnS	Zinc sulfide
Potassium	Bromine	KBr	Potassium bromide
Calcium	Carbon	CaC_2	Calcium carbide
Sodium	Chlorine	$NaCl$	Sodium chloride
Hydrogen	Sulfur	H_2S	Hydrogen sulfide
Hydrogen	Chlorine	HCl	Hydrogen chloride
Lead	Iodine	PbI_2	Lead iodide
Ammonium	Chlorine	NH_4Cl	Ammonium chloride

97. In some cases an element can unite with another element to form more than one compound. There is a system of naming these substances so that they can be distinguished from one another. For example, iron may unite with chlorine to form two different kinds of chlorides, depending on the number of atoms of chlorine joined to 1 atom of iron, or depending on the valence of the iron in each chloride. The chloride $FeCl_2$, in which iron is divalent and in which but 2 atoms of chlorine are present to every atom of iron, is called ferrous chloride. The substance $FeCl_3$, in which iron is trivalent and in which 3 atoms of chlorine are present to every atom of iron, is called ferric chloride. In other words, the ending *ous* is added to

the name of the positive element, usually a metal, to show that it is present with the lower valence and that it is combined with a lower amount of negative element than in the compound in which the positive element has the ending *ic*.

Following are a few examples showing the use of the endings *ous* and *ic*:

Cu_2O, cupr*ous* oxide	CuO, cupr*ic* oxide
$SnCl_2$, stann*ous* chloride	$SnCl_4$, stann*ic* chloride
CrO, chrom*ous* oxide	Cr_2O_3, chrom*ic* oxide

98. Elements Uniting in More Than One Proportion. Another method, in which a prefix is attached to the negative element, is used to distinguish binary compounds formed by elements uniting in more than one proportion. The prefixes *mono; di,* or *bi; tri; tetra;* and *penta* are applied to show whether the number of atoms of the negative element present is, respectively, one, two, three, four, or five. Following are a few examples of this system:

CO, carbon monoxide	CO_2, carbon dioxide
SO_2, sulfur dioxide	SO_3, sulfur trioxide
PCl_3, phosphorus trichloride	PCl_5, phosphorus pentachloride

The prefix *per* is sometimes applied to the negative element to show that the greatest possible number of atoms of the negative element is present, thus:

H_2O_2, hydrogen peroxide	Na_2O_2, sodium peroxide
BaO_2, barium peroxide	

No compound of hydrogen and oxygen has a greater proportion of the negative element than has hydrogen peroxide. The same fact holds true for all other binary compounds in which the prefix *per* appears.

The following oxides of nitrogen are excellent examples of substances that follow the nomenclature explained in the foregoing:

N_2O, nitrous oxide
NO, nitric oxide
N_2O_3, nitrogen trioxide
NO_2 or N_2O_4, nitrogen dioxide or nitrogen tetroxide
N_2O_5, nitrogen pentoxide

The prefix *sesqui* is sometimes added to the negative element to denote that there is present $1\frac{1}{2}$ atoms of the negative element to every atom of the positive element. For example, Fe_2O_3, ferric oxide, is sometimes called iron sesquioxide.

NOMENCLATURE OF ACIDS AND SALTS BELONGING TO ONE SERIES

99. Series of Acids.—The system of naming acids and salts should be studied carefully, for it is used in every phase of chemistry. Some acids so far considered have names ending in *ic*. Others end in *ic,* but have in addition the prefix *hydro;* others, again, have different prefixes and suffixes. However, each one of them is based on the same system of naming substances.

An acid is generally named by referring to its characteristic element. It should be borne in mind that in an acid, neither hydrogen nor oxygen can be considered as a characteristic element, for the former is found in all and the latter in most acids. For example, H_2SO_4 is called sulfuric acid, and HNO_3, nitric acid, being named with reference to the respective characteristic elements sulfur and nitrogen.

The principal acid of a series takes the ending *ic,* as $HClO_3$, chloric acid; the next acid in the series, containing less oxygen than the principal acid, takes the ending *ous,* as $HClO_2$, chlorous acid; the acid containing even less oxygen than the acid ending in *ous,* takes the ending *ous* and the prefix *hypo,* as $HClO$, hypochlorous acid; the acid of the series containing hydrogen and the characteristic element only, takes the ending *ic* and the prefix *hydro,* as HCl, hydrochloric acid. If there is an acid in the series containing more oxygen than the principal acid, it has the ending *ic* and the prefix *per,* as $HClO_4$, perchloric acid.

Following is a summary of the formulas and names of the acids of chlorine:

HCl, *hydro*chlor*ic* acid
$HClO$, *hypo*chlor*ous* acid
$HClO_2$, chlor*ous* acid
$HClO_3$, chlor*ic* acid
$HClO_4$, *per*chlor*ic* acid

100. The foregoing names have been used for years. There is also a set of names based on scientific grounds, that can properly be used to designate the various chlorine acids, as follows:

HCl, hydrogen chlor*ide*
$HClO$, hydrogen *hypo*chlor*ite*
$HClO_2$, hydrogen chlor*ite*
$HClO_3$, hydrogen chlor*ate*
$HClO_4$, hydrogen *per*chlor*ate*

101. Series of Salts.—The names of salts depend on the names of the acids from which they are formed. A salt formed by a metal and the negative radical of an oxyacid ending in *ic* has the ending *ate;* and a salt consisting of a metal and the negative radical of an oxyacid ending in *ous* has the ending *ite*. The prefix appearing in the name of the acid is retained in the name of the salt. As has been explained before, names of salts consisting of but two elements end in *ide,* even though the acid has the prefix *hydro* and the suffix *ic*. Names of salts of chlorine acids receive the same endings as the acids, as shown in the foregoing explanations, excepting that the name of the metal in the salt is used instead of hydrogen. Following are the names of different salts formed when potassium hydroxide reacts with the acids of chlorine:

KCl, potassium chlor*ide*; formed from HCl, hydro*chlor*ic acid.

$KClO$, potassium *hypo*chlor*ite;* formed from $HClO$, *hypo*chlor*ous* acid or hydrogen *hypo*chlor*ite*.

$KClO_2$, potassium chlor*ite*; formed from $HClO_2$, chlor*ous* acid or hydrogen chlor*ite*.

$KClO_3$, potassium chlor*ate*; formed from $HClO_3$, chlor*ic* acid or hydrogen chlor*ate*.

$KClO_4$, potassium *per*chlor*ate*; formed from $HClO_4$, *per*chlor*ic* acid or hydrogen *per*chlor*ate*.

102. Salts containing positive elements that exhibit more than one valence are named in much the same way as binary compounds containing positive elements that have more than one valence. The name of the positive element, when it has

the lower valence, takes the ending *ous*, and when it has the higher valence, takes the ending *ic*. Thus:

$SnSO_4$, stannous sulfate
$Sn(SO_4)_2$, stannic sulfate

Each salt is a sulfate, for each one of them contains the negative radical, SO_4, of sulfur*ic* acid. The salt in which tin is divalent is called stann*ous* sulfate, and the one in which it is tetravalent is called stann*ic* sulfate.

103. Formulas and Names of Some Common Acids and Salts.—The following examples of acids and their salts are given in order to illustrate further the general application of the nomenclature explained in the preceding pages:

Acids	Sodium Salts
Hydrochlor*ic* acid, HCl	Sodium chlor*ide*, $NaCl$
Sulfur*ous* acid, H_2SO_3	Sodium sulf*ite*, Na_2SO_3
Sulfur*ic* acid, H_2SO_4	Sodium sulf*ate*, Na_2SO_4
Nitr*ous* acid, HNO_2	Sodium nitr*ite*, $NaNO_2$
Nitr*ic* acid, HNO_3	Sodium nitr*ate*, $NaNO_3$
Phosphor*ous* acid, H_3PO_3	Sodium phosph*ite*, Na_2HPO_3
Phosphor*ic* acid, H_3PO_4	Sodium phosph*ate*, Na_3PO_4

Miscellaneous Salts and Compounds

Cupr*ous* oxide, Cu_2O	Sodium *per*chlor*ate*, $NaClO_4$
Cupr*ic* oxide, CuO	Stann*ous* chlor*ide*, $SnCl_2$
Mercur*ic* chloride, $HgCl_2$	Calcium chlor*ate*, $Ca(ClO_3)_2$
Mercur*ous* chlor*ide*, $HgCl$	Sulfur *di*ox*ide*, SO_2
Ferr*ous* sulf*ate*, $FeSO_4$	Sulfur *tri*ox*ide*, SO_3
Ferr*ic* sulf*ate*, $Fe_2(SO_4)_3$	Cupr*ous* sulf*ide*, Cu_2S
Ferr*ous* sulf*ite*, $FeSO_3$	Cupr*ic* sulf*ide*, CuS
Ferr*ic* sulf*ite*, $Fe_2(SO_3)_3$	

Sodium *bi*sulf*ate*, acid sodium sulf*ate*, or sodium hydrogen sulf*ate*, $NaHSO_4$

Acid sodium carbon*ate*, sodium hydrogen carbon*ate*, or sodium *bi*carbon*ate*, $NaHCO_3$

Potassium cyan*ide*, KCN

Calcium *hypo*chlor*ite*, $Ca(ClO)_2$

INORGANIC CHEMISTRY

Serial 5560A (PART 1) Edition 1

EXAMINATION QUESTIONS

Notice to Students.—*Study the Instruction Paper thoroughly before you attempt to answer these questions.* **Read each question carefully and be sure you understand it;** *then write the best answer you can. When your answers are completed, examine them closely, correct all the errors you can find, and* **see that every question is answered;** *then mail your work to us.*

(1) (*a*) Define chemistry. (*b*) What is the difference between organic and inorganic chemistry?

(2) Balance the following equations:

$$H_2+Fe_3O_4 \longrightarrow Fe+H_2O$$
$$Na_2O_2+H_2O \longrightarrow NaOH+O_2$$
$$KClO_3+ \xrightarrow{\Delta} KCl+O_2$$
$$P_2O_5+H_2O \longrightarrow H_3PO_4$$
$$N_2+H_2 \longrightarrow NH_3$$

(3) Define (*a*) chemical compound; (*b*) element; (*c*) atom; (*d*) molecule.

(4) (*a*) What temperature scale is commonly used by chemists? (*b*) What are the freezing and boiling points of water on the centigrade and Fahrenheit scales? (*c*) What is meant by *absolute zero*?

(5) (*a*) Convert the following temperatures: 95° C. to F.; 95° F. to C.; −13° F. to C.; and −10° C. to F. (*b*) Express −100° C. and 100° C. in terms of absolute temperature.

(6) State the difference between the weight and the mass of a body.

(7) (*a*) Define matter. (*b*) What are the three physical states of matter? (*c*) Give an example of each physical state of matter.

(8) What are the five types of chemical reactions? Give an example of each.

(9) What conditions must be satisfied by a correctly written chemical equation?

(10) (*a*) State Charles' law. (*b*) State Boyle's law.

(11) A gas occupies a volume of 150 liters at a pressure of 720 mm. and a temperature of 20° C. What volume will the gas occupy at standard conditions (760 mm. and 0° C.)?

(12) A gas confined over water occupies a volume of 425 ml. at a temperature of 29° C. and 760 mm. pressure. What will be the volume of the dry gas under standard conditions?

(13) (*a*) How many grams are there in one kilogram? (*b*) How many milliliters are there in one liter? (*c*) What is the weight of one milliliter of water at 4° C.?

(14) Discuss chemical and physical changes. Show by examples the difference between these two kinds of changes.

(15) Define the following and give an example of each: (*a*) Normal salt; (*b*) acid salt; (*c*) basic salt; (*d*) double salt.

NOTE: Show how you obtain answers to questions that involve the solution of problems.

Mail your work on this lesson as soon as you have finished it and looked it over carefully. DO NOT HOLD IT until another lesson is ready.

Memo:

pls check formula $Ca(OH)_2$

memorize all formulas & equations

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